TEACHERS GUIDE WITH ANSWERS

INTRODUCTORY CHEMISTRY AND PHYSICS

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PREFACE

This manual has been prepared for teachers who are using Introductory Chemistry and Physics by Pickard and Radomsky.

In Section I are a few suggestions on teaching this course. This section includes:

(a) The Bohr atom - its usefulness and its limitations.

(b) Nomenclature - some simple rules that should be taught in conjunction with Chapter 5.

(c) Problem solving - requisites of problem solving.

Section II deals with laboratory work. This section includes:

(a) Importance of laboratory work.

(b) Care and maintenance of the laboratory.

(c) Preparing solutions for use in bench reagent bottles.

(d) Laboratory equipment required for this course.

(e) Laboratory exercises.

Section III contains answers to the questions and solutions to the problems in the textbook.

Section IV contains review exercises.

Section II (b), (c), and (d) were taken from the Curriculum Guide for Senior High School Science for the Province of Albertâ (1962). We wish to express our gratitude to the officials of the Department of Education for allowing us to use this material.

H. J. K.

S. W. R.

TEACHERS GUIDE WITH ANSWERS

INTRODUCTORY CHEMISTRY AND PHYSICS

SECTION I

METHOD

It is not the purpose of this manual to state what method or methods a teacher should use. There is no single method for the successful teaching of science. The methods used must be adapted to the type of class, needs of the students, nature of content and the personality of the teacher. A good teacher shows a high degree of versatility in using a variety of methods.

The textbook and the science curriculum are adapted to promote the aims and objectives of science teaching. The science teacher is expected to create an atmosphere conducive to the inculcation and appreciation of these aims and objectives. Books, equipment, building and curricula are only aids to better instruction.

Every good teacher uses analogies to develop concepts and principles. Analogies can be very effective but misleading if accepted uncritically, Teachers should therefore apply them in such a way as to make clear their limitations.

Teachers are prone to telling the students too much - to "spoonfeed" them. Students should be given every opportunity to think for themselves. The role of the science teacher is that of a guide or adviser; his major responsibility is to suggest, not to tell, For example, if a demonstration is used as a method for introducing a concept or a principle, this should be done with as few comments as possible. The students should record the observations and draw their own conclusions to be discussed in class later.

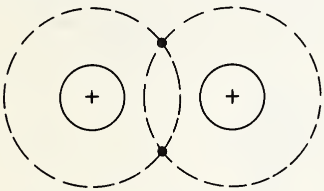
ATOMIC STRUCTURE

As the text is largely introductory in intent, the simplified Bohr-Rutherford atom is introduced, since this model serves to explain most of the general principles included.

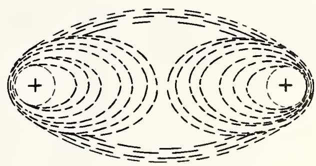
Originally Bohr proposed that electrons move in curved orbits about the nucleus and that as this motion was three-dimensional, the path of the electron may be described as a spherical shell. As the total energy of the electron is quantized, i.e. limited to certain values, all the electrons in a given shell are characterized by one energy level, the shell closest to the nucleus corresponding to the lowest energy level, and the shells farther away corresponding to higher energy levels. Although this oversimplification is no longer acceptable, the terms "shell" and "orbit" are still sometimes used to denote energy levels. The standard designations of the energy levels as K, L, M, ..., are therefore introduced.

The text discusses at some length only the first twenty elements. By means of the simplified diagrams, the general characteristics of these elements with respect to electronegativity, electrovalence, and chemical activity are introduced to the student, possibly for the first time. It should be noted that electronic configurations alone do not account for all chemical properties of the elements and that the arrangement of the dots representing electrons has no significance as far as the actual positions of the electrons are concerned. The electrons are, however, represented in pairs according to the Pauli exclusion principle although the treatment of electron configurations with respect to suborbitals and the electron probability distribution are beyond the scope of an introductory course.

The limitations of the Bohr atom should be pointed out to the students, and wherever possible the electron cloud configuration should be introduced and discussed briefly.



Simplified Bohr atom representation



Electron cloud configuration

A molecule of hydrogen illustrating covalence

NOMENCLATURE IN CHEMISTRY

The naming of compounds should be taught in conjunction with Chapter 5. Once the students have mastered the valence concept and how it is used in the writing of formulae, some of the rules for naming compounds should be presented to them.

Lavoisier began the system of nomenclature used today. This system of naming compounds is designed to supply information concerning the elements contained in the compound.

BINARY COMPOUNDS

Binary compounds consist of two elements only. The metallic or the positive part is written first and then the nonmetallic or the negative part. e.g.

$NaCl$ - sodium chlor<u>ide</u>

BaO - barium ox<u>ide</u>

H_2S - hydrogen sulf<u>ide</u>

The name of a binary compound ends in -<u>ide</u>.

<u>OXIDES AND PEROXIDES</u>

An oxide is a binary compound containing the element oxygen. e.g.

Na_2O - sodium ox<u>ide</u>

BaO - barium ox<u>ide</u>

Al_2O_3- aluminum ox<u>ide</u>

Note that the above examples conform to the valence rules.

There are some compounds and oxides which do not conform to valence rules. They possess one more atom of oxygen per molecule than the normal oxide. Such oxides are designated by the prefix per-. e.g.

Na_2O_2- sodium <u>per</u>ox<u>ide</u>

H_2O_2 - hydrogen <u>per</u>ox<u>ide</u>

BaO_2 - barium <u>per</u>ox<u>ide</u>

<u>VARIABLE VALENCE FOR METALS</u>

When a metal has two different valences, the suffixes -<u>ic</u> and -<u>ous</u> are used.

The suffix -<u>ic</u> denotes the <u>higher</u> valence.

The suffix -<u>ous</u> denotes the <u>lower</u> valence.

e.g. $CuBr_2$- cupr<u>ic</u> brom<u>ide</u>

$CuBr$ - cupr<u>ous</u> brom<u>ide</u>

FeO - ferr<u>ous</u> ox<u>ide</u>

Fe_2O_3- ferr<u>ic</u> ox<u>ide</u>

<u>VARIABLE VALENCE FOR NON-METALS</u>

Sometimes non-metals act as metals in the formation of compounds; that is, they combine with other non-metals. In the naming of such compounds the suffixes are not usually used. Prefixes such as mon- or mono-, di-, tri-, tetra-, pent- or penta-, etc. are used before the name of the second element to indicate the number of atoms of the second element. e.g.

CO	-	carbon monoxide
CO_2	-	carbon dioxide
SO_3	-	sulfur trioxide
CCl_4	-	carbon tetrachloride
P_2O_5	-	phosphorus pentoxide

BINARY ACIDS

Binary acids are made up of two elements, one of which is hydrogen. Some binary hydrogen compounds (gases) when dissolved in water react with water to form binary acids. The formula of the binary acid is the same as that of its gas.

The name of a binary acid begins with the prefix hydro- and ends with the suffix -ic.

COMPOUND	ACID
HF - hydrogen fluoride	HF - hydrofluoric acid
HCl - hydrogen chloride	HCl - hydrochloric acid
HBr - hydrogen bromide	HBr - hydrobromic acid
HI - hydrogen iodide	HI - hydroiodic acid
H_2S - hydrogen sulfide	H_2S - hydrosulfuric acid

OXY-ACIDS OR TERNARY ACIDS

Oxy-acids consist of three elements, two of which are hydrogen and oxygen. The standard oxy-acid ends with -ic, but often the number of atoms of oxygen in a molecule of acid varies.

If a molecule of an oxy-acid contains one more atom of oxygen than the standard -ic acid, it is designated by the prefix per- and the suffix -ic.

If the molecule contains one atom of oxygen less than the standard -ic acid, it is designated by the suffix -ous.

If the molecule contains two atoms of oxygen less than the -ic acid, it is designated by the prefix hypo- and the suffix -ous.

$HClO_4$	-	perchloric acid
$HClO_3$	-	chloric acid
$HClO_2$	-	chlorous acid
HClO	-	hypochlorous acid

BASES

The molecule of a base consists of a metal or metallic radical and the hydroxyl radical. The hydroxyl radical is non-metallic in nature (electrovalence of -1) and therefore appears at the end.

Bases are named hydroxides.

NaOH	-	sodium hydroxide
NH_4OH	-	ammonium hydroxide
$Al(OH)_3$	-	aluminum hydroxide

SALTS

There are several ways in which salts are formed. We may consider salt as the product formed when the replaceable hydrogen of an acid is displaced by a metal or a metallic radical. Binary salts are formed from binary acids. Like all binary compounds, binary salts end in -ide.

BINARY ACID			BINARY SALT		
HCl	-	hydrochloric acid	NaCl	-	sodium chloride
HBr	-	hydrobromic acid	KBr	-	potassium bromide
HF	-	hydrofluoric acid	CaF_2	-	calcium fluoride
H_2S	-	hydrosulfuric acid	CuS	-	cupric sulfide

Oxy-acids form oxy-salts. The name of the salt has a definite relationship to the name of its corresponding acid.

A per- -ic acid forms a per- -ate salt.

An -ic acid forms an -ate salt.

An -ous acid forms an -ite salt.

An hypo- -ous acid forms an hypo- -ite salt.

ACID			SALT		
$HClO_4$	-	perchloric acid	$NaClO_4$	-	sodium perchlorate
$HClO_3$	-	chloric acid	$NaClO_3$	-	sodium chlorate
$HClO_2$	-	chlorous acid	$NaClO_2$	-	sodium chlorite
HClO	-	hypochlorous acid	NaClO	-	sodium hypochlorite

NORMAL SALTS AND ACID SALTS

Normal salts are formed when all the replaceable hydrogen of an acid is displaced by a metal or a metallic radical.

H_2SO_4 - sulfuric acid

Na_2SO_4 - sodium sulfate (normal salt)

ACID SALTS are formed when only part of the replacable hydrogen of an acid is displaced by a metal or a metallic radical. Acid salts are named by using the prefix bi- before the nonmetallic radical.

e.g.

$NaHSO_4$ - Sodium bisulfate

$KHCO_3$ - potassium bicarbonate

NH_4HSO_4 - ammonium bisulfate

PROBLEM-SOLVING

This course stresses the importance of problem-solving. Problem-solving promotes the development of critical and analytical thinking and gives the student a better understanding of the concepts and principles of science.

Four essentials to problem-solving are:

(1) The ability to read and interpret problems. This ability may be improved if the student is encouraged to draw simple line diagrams illustrating his interpretation of the problem.

(2) A thorough knowledge of definitions. A definition should not be treated as something to be memorized and reproduced on paper; rather, as a "crutch" to help solve problems.

(3) Use of complete statements. "Plugging-in" of formulae in solving problems should be discouraged. If a student uses statements, the teacher can follow his line of reasoning and spot his difficulties with ease.

(4) Explicit use of units of measurement.

Students should be encouraged to use units of measurement explicitly and to treat them as algebraic expressions. Students may find this type of work awkward at first, but with continued use they will develop an appreciation for the importance of using units of measurement correctly. This eliminates guesswork. They are no longer "pinning tails to donkeys."

Students should be encouraged to write large and small numbers as powers of ten. This is easy and compact. Any number can be written as the product of a number between one and ten and a number which is a power of ten.

e.g.

$39700 = 3.97 \times 10^4$

$0.126 = 1.26 \times 10^{-1}$

$93{,}000{,}000 = 9.3 \times 10^7$

The exponent, at the upper right of the 10, tells how many zeros in the power of ten, and the minus sign is used when dealing with fractions.

SECTION II

IMPORTANCE OF LABORATORY WORK

This course is based on laboratory work. The laboratory is commonly regarded as the heart of science. All science teaching procedures should be closely related to laboratory work. Other procedures should lead, in general, to laboratory experience or result from it. The laboratory should provide basic experiences which are essential to the promotion of the natural scientific interests of each individual. Laboratory work develops critical thinking.

This course is based on the premise that it may be used in overcrowded classrooms, schools with limited laboratory facilities, and classrooms filled with students who are not primarily interested in becoming scientists. To this end, the laboratory exercises have been prepared with so much care outlining the procedures for students to follow that they may well be labelled "cook books". They appear to be designed primarily to illustrate and to verify. The use of a laboratory as a teaching method is left entirely to the discretion of the teacher. It is up to the teacher to adapt laboratory work to the needs of the class. However, it should be pointed out that a directed laboratory may be used to advantage in science teaching.

(a) It is time-saving and time is essential to good teaching.

(b) It provides for illustrations of principles and their applications.

(c) It provides the means to verify facts, laws and generalizations.

(d) It contributes to the students' knowledge and understanding of facts, principles, concepts, and generalizations.

(e) It contributes to the development of skills, habits and attitudes.

(f) It gives the teacher a feeling of security, for all the necessary apparatus and supplies are anticipated.

(g) It gives the student direction leading to a feeling of success and accomplishment.

(h) It is, in itself, a safety precaution.

It is not argued here that a student should not be encouraged to truly experiment - to find his own problem and solve it in the laboratory. On the contrary, this aspect of laboratory work should be promoted and encouraged by the teacher as much as possible. This kind of laboratory work may be promoted in schools through the organization of science clubs. Students who are genuinely interested in science should join the science club and be encouraged to pursue the topics of their interest.

Time is required for laboratory work. This course is based on the premise that each student will spend at least 20% of his class time in the laboratory. It takes time to set up apparatus, perform the several connected steps, and to write down the observations.

Inadequate time allowance has led many teachers to adopt short cuts in the laboratory. Experiments are often simplified and shortened, removing any substantial thinking on the part of the student. Apparatus, including glass tubing, is all set up so that it goes together like a mechanical toy. Such devices are far removed from the work of a scientist and are but a poor substitute for the teaching of science. For example, every student should learn to bend and draw out glass tubing, and to fire-polish the sharp edges.

Teachers working on 35-40 minute class periods should use double periods for laboratory work, thus giving the student at least one hour of continuous work.

It should also be emphasized that the proper timing of laboratory exercises is essential to effective science teaching.

CARE AND MAINTENANCE OF THE SCIENCE LABORATORY

1. Chemicals and equipment should be stored in a neat and orderly manner. Chemicals can well be arranged alphabetically. All equipment should be cleaned before being put away.

2. All apparatus should be kept in good repair. The teacher should be responsible for minor repairs. For major repairs equipment can usually be returned to the supply house from which it was purchased.

3. Damaged equipment (chipped glassware) should be discarded.

4. A yearly inventory should be taken each spring to determine basic needs. **DO THIS NOW.** The science teacher is responsible for seeing that chemicals and equipment are ordered and available as required. Purchase of some of the major items of equipment may be spread over a term of several years.

5. A resourceful teacher will find many ways of improvising equipment for experimental work.

6. Care should be taken in storing chemicals, e.g., the kerosene level on sodium should be checked, water level on phosphorus, etc. Acids should be stored in a closed to prevent fumes from reacting with metal equipment.

7. The laboratory should be locked when a teacher is not present for supervision.

8. Students should be taught how to use and care for laboratory equipment.

9. A fire extinguisher should be provided in each science laboratory and should be easily accessible. Students should be instructed in the proper use of the extinguisher.

10. A first-aid kit should be available.

11. Baking soda and a mild solution of acetic acid should be kept on hand for treating acid and alkali burns. Students should be warned about treating such burns immediately.

12. Students should wear laboratory coats or aprons when working in the laboratory. (Dad's old shirt will do.)

PREPARING SOLUTIONS FOR USE IN BENCH REAGENT BOTTLES.

Methyl orange indicator: Dissolve one gram of methyl orange in 1000 ml. of water.

Phenolphthalein indicator: Dissolve one gram of phenolphthalein in 700 ml. of methyl alcohol and dilute to 1000 ml. with water.

Ammonium hydroxide 6N: Add 400 ml. of concentrated ammonium hydroxide to 600 ml. of water.

Barium chloride 1N: Dissolve 122 grams of $BaCl_2 \cdot 2H_2O$ in 500 ml. of water and after the salt has dissolved, and sufficient water to make 1000 ml. of solution.

Hydrochloric acid 6N: Add 500 ml. of concentrated hydrochloric acid to 500 ml. of water.

Nitric acid 6N: Add 380 ml. of concentrated nitric acid to 620 ml. of water.

Silver nitrate 0.1N: A 0.1 normal solution is suitable for the purposes of this course and may be made by adding 17 grams of silver nitrate to sufficient water to make 1000 ml. of solution.

Sodium chloride 1N: 58.5 grams of sodium chloride plus enough water to make 1000 ml. of solution.

Sodium hydroxide 2.5N: 100 grams of sodium hydroxide to enough water to make 1000 ml. of solution.

Sulfuric acid 6N: Add 200 ml. of concentrated sulfuric acid to enough water to make 1000 cc. of solution.

Caution. When mixing acid and water, always add the acid to the water slowly, stirring the solution.

LABORATORY EQUIPMENT FOR SCIENCE

Student Experimental Work

It is suggested that students should work in groups of two. Each group should be equipped with the following apparatus, which is considered to be a minimum. This apparatus should be stored at the laboratory station at which the group works.

CHEMISTRY EQUIPMENT SCIENCE 10

APPARATUS (for a group of two)

Beakers 2 - 150 ml. capacity
        1 - 250 ml. capacity
1 mortar and pestle (small size)
1 Florence flask (500 cc.)
1 deflagrating spoon
1 doz. test tubes (medium size, Pyrex)
2 test-tube clamps
3 crucibles and covers (size '0')
2 four-inch wire gauze (asbestos)
1 crucible tongs
2 evaporating dishes
1 Bunsen burner
1 Erlenmeyer flask (500 ml.)
2 rubber tube clamps
1 graduated cylinder (50 ml.)
1 test-tube brush
1 set of ring stand and rings
1 funnel (small)
2 utility clamps
1 wire triangle (clay guards)
3 wide-mouth gas bottles
1 four-inch triangular file
1 wing top for Bunsen burner
1 spatula (small)
1 asbestos mat (12" square)
1 thermometer (-12 to 110^{o}C)
1 pneumatic trough
6 reagent bottles (4 oz. size - named for chemical reagents, glass stoppers)

REAGENTS AND MATERIALS REQUIRED (based on a group of twenty students)

	Quantity
Alcohol, Methyl (local purchase)	1 gal.
Aluminum, sheet	1 sq. ft. 1/32" thick
Ammonium hydroxide	4 lbs.
Barium Chloride, cyrstals CP	1 lb.
Barium Chloride, Tech	1 lb.
Calcium Carbonate (marble chips)	5 lbs.
Calcium Chloride, Gran. USP	5 lbs.
Copper, sheet (for Strips)	1 lb. 1/32" thick
Copper Sulphate, cyrstals CP	1 lb.
Cupric Chloride, anhydrous	1/4 lb.
Hydrochloric Acid, concentrated Sp. G. 1.19	6 lbs.
Iron filings	1 lb.
Iron nails (obtain locally)	
Lead, sheet	1 sq. ft. 1/16" thick
Lemon juice (local purchase)	
Litmus	4 ozs.
Manganese Dioxide CP	3 lbs.

	Quantity
Magnesium ribbon	1 oz.
Mercuric Oxide CP.	1/4 lb.
Methyl Orange.	1 oz.
Nitric Acid USP.	7 lbs.
Phenolphthalein.	4 ozs.
Potassium Chloride	1 lb.
Potassium Chlorate CP.	2 lbs.
Salt (not iodized) Granular USP	
Sand (local purchase).	5 lbs.
Silver Nitrate	
Sodium Hydroxide, pellets USP.	1 lb.
Sodium, metal.	1/4 lb.
Sodium Sulphate, crystals CP	1 lb.
Solder, spool (local purchase)	
Sugar, cubes (local purchase)	
Sulfur (flowers of S.) powdered.	1 lb.
Sulfuric Acid, SP. G. 1.83	9 lbs.
Vinegar (local purchase)	
Zinc (mossy) (low in As, Fe and Pb).	1 lb.

CP - Highest degree of purity. USP - Less Pure. Tech - Contains some impurities.

PHYSICS EQUIPMENT SCIENCE 10

Minimum Apparatus for a Class of 20 - 24 students

1 Brownian movement apparatus
*1 set of Pascal Vases with base attachment
1 set of capillary tubes
1 barometer
*1 set of Magdeburg hemispheres
1 exhaust pump with base and bell jar
1 Torricellian tube
1 open-tube manometer
1 Bourdon pressure gauge
1 lift pump model
1 Boyle's law apparatus
12 thistle tubes
1 Venturi tube
3 balance scales - (Harvard Trip)
3 sets of weights
5 glass U-tubes
5 lb. bottle methyl alcohol
5 lb. bottle carbon tetrachloride
5 lbs. mercury
5 thermometers (centigrade)
5 calorimeters
2 coefficient of linear expansion apparatus with spherometer and rods
2 adhesion disks
1 adjustable thermostat
4 dilatometers
1 ball and ring (hand form)

1 bi-metallic strip
2 Charles' Law tubes (Waterman form)
2 boilers or steam generators
2 sets specific heat specimens
4 water traps (for heat of vaporization)
4 meter sticks
24 beakers
glass tubing
rubber delivery tube
rubber stoppers (6 lb. assorted)
4 graduated cylinders 100 cc.
2 glass tube cutters
retort stands
clamps and rings
12 Bunsen burners or propane burners
2 sets of tuning forks (unmounted)
1 set of tuning forks (mounted)
2 rubber hammers
*1 Kundt's tube
*1 pint cork dust
1 steel rod
1 brass rod
1 coil spring (longitudinal vibrators)
1 electric bell in vacuo
1 rotator (hand driven, worm drive)
1 Savart's toothed wheels
1 Carova's disc
1 combined siren and color disc
1 sonometer with strings of varied densities and diameters
1 set of collision balls
2 sets of resonance tubes

*Optional equipment

EQUIPMENT REQUIRED BY THE TEACHER FOR GROUP OF ANY SIZE

Balances, Chemical:

There should be at least one chemical balance capable of weighing to 0.01 grams. In addition, a rough balance for making dilute and bench solutions is a handy piece of equipment. See Physics List.

Glass Tubing:

A supply of glass tubing which can be issued to students for glass work should be kept in stock. A popular size and type is the soft-glass tube with the 5 or 6 mm. outside diameter. Fifteen to twenty pounds should be kept in stock for a class of twenty students. It can be used for delivery tubes for gas production, making of wash bottles and exercise in glass work. Do not buy the Pyrex tube as it requires too much heat to soften it to the stage where it can be worked.

Rubber Tubing:

A supply for rubber tubing is necessary for connecting gas burners, delivery tubes, etc. The usual rubber tubing for connecting gas burners has an inside bore of 1/4 of an inch. Delivery tubes made of 6 mm. glass may be connected by using rubber tubing with a 3/16-inch bore (inside diameter).

Rubber Stoppers:

A quantity of rubber stoppers, solid, one-hole and two-hole are needed. The most useful sizes are No. 2, 3 and 4. In some cases it is well to have a few No. 5 stoppers as well.

Beakers:

An assortment.

Gas Generators:

The Kipp generator for gases is practically a "must" for any school laboratory producing gas. Great care should be exercised in any work with hydrogen gas. Hydrogen sulfide cartridges are available.

Graduated Cylinders:

One or two graduated cylinders of one liter capacity are necessary in the making of stock solutions.

Filter Paper:

Several packages, about 9 cm. diameter, are needed.

Litmus Paper:

Red and blue, obtainable in books of 100 test strips each, are needed. About twenty books of each give a suitable stock.

Distilled Water, etc.

A supply of distilled water should be available. It is necessary for making solutions of silver nitrate especially if tap water contains chlorine or any of the chlorides. It should be used for all stock solutions. Obtain locally. Cloths, soap, towels, matches, splints, etc. are always necessary.

Crookes' Tube and Induction Coil, Pith Balls, Rods, Fur and Silk Cloth:

This equipment is desirable to demonstrate the lessons on the electron.

EXPERIMENTS

In carrying out the following experiments, at least a forty-minute laboratory period is required. To this view, in some cases, shorter experiments dealing with the same general topic have been incorporated into one experiment. The text describes many other experiments which may also be combined on the basis of the time available.

Experiments 1, 3, 4 and 8 are quantitative in nature. An error calculation and a critical analysis of the results are desirable in each case.

Experiments 5 and 9 should be demonstrated by the teacher because of the dangers involved. Experiment 5 may be used to introduce Chapter 9 and experiment 9 to introduce Chapter 8.

EXPERIMENT 1.

MEASUREMENT

Object:

To make measurements of length, volume and mass using the metric system and to become familiar with relationships existing between them.

Apparatus:

A ruler graduated in millimeters or a vernier caliper, a 100 milliliter graduated cylinder, a platform balance, a small cylindrical can.

Procedure:

(1) By means of the vernier determine the height and diameter of the can.

(2) Calculate the volume of the container.

(3) Weigh the empty container (in grams).

(4) Fill the graduated cylinder with water and determine the volume of water required to fill the can (always read the bottom of the meniscus).

(5) Weigh the can and water.

(6) Determine the mass of water.

Observations:

Height of can (cm) =

Diameter of can (cm) =

Weight of can empty =

Weight of can and water =

Calculations:

Volume of can in cc.=
Volume of can in ml.=
Mass of water in gm.=
Density of water in gm/cc.=

Conclusion:

1. Assuming the density of water to be 1 gm/cc. calculate the per cent error.

Note: Per cent error=

$$\frac{\text{difference between correct value and obtained value}}{\text{obtained value}} \times 100\%$$

2. State sources of error.

EXPERIMENT 2.

PREPARATION OF A COMPOUND

Object:

To combine two elements chemically by heating them together.

Apparatus and Materials:

Porcelain crucible ring stand, clay triangle, Bunsen burner, sulfur, copper foil, pestle, crucible tongs.

Diagram:

Draw and label a diagram showing the arrangement of apparatus.

Procedure:

(1) Examine the starting materials and list their apparent properties.

(2) Place a thin layer of sulfur at the bottom of the crucible.

(3) Place the piece of copper on the sulfur layer.

(4) Place a thin layer of sulfur over the copper, leaving some copper exposed.

(5) Arrange the apparatus in the fume cabinet or on a bench near the window.

(6) Begin heating the crucible with a cool Bunsen flame, continuing until the sulfur is melted.

(7) Note all changes taking place in the starting materials (Do not become alarmed if the contents of the crucible catch fire. Do not attempt to blow the fire out. Remove the Bunsen burner.)

(8) Increase the temperature of the flame and continue heating until the reaction is complete.

(9) Allow the contents to cool; note all properties; when cool, grind the remains with a pestle.

Observation:

Properties of:

COPPER	SULFUR	COPPER SULFIDE

Conclusion:

(1) Decide, giving the reasons, whether the remains in the crucible were a mixture of copper and sulfur, copper alone, sulfur alone, or a new substance.

(2) Write the chemical equation describing the reaction.

EXPERIMENT 3.

THE DETERMINATION OF PER CENT WATER OF HYDRATION IN BLUESTONE

Object:

To determine the per cent water of hydration in bluestone crystals.

Apparatus and Materials:

Crucible, Bunsen burner, bluestone crystals, tongs, balance.

Theory:

Water of hydration forms a definite proportion by weight of many crystals. The number of water molecules present in the hydrated crystal of a particular compound is the same for all samples although it varies for crystals of different substances.

Procedure:

(1) Clean and weigh a porcelain crucible and place two or three bluestone crystals in it. Note the properties of bluestone (color, appearance, etc.).

(2) Weigh the crucible and bluestone to the nearest centrigram.

(3) Place the crucible on the clay triangle and support it on the ring stand. Begin heating the crucible with a relatively cool Bunsen flame.

(4) After a few minutes, increase the intensity of the flame. Note the changes taking place. Continue heating until the change is complete and then heat strongly for 5 minutes.

(5) Cool the crucible and its contents and when cool, weigh it.

(6) Heat the crucible and its contents very strongly for 3 minutes, cool and weigh as before. If the weight has not changed, assume that the reaction has reached completion. If the weight has altered, heat again, continuing this procedure until two consecutive weighings produce the same results. This process is called "heating to a constant weight".

Observations and Calculations:

1. Record any changes in appearance.

2. Record the weights as follows:
   wt. of crucible =
   wt. of crucible + bluestone =
   wt. of bluestone =
   wt. after first heating =
   wt. after second heating =
   change in weight =

Assume that the decrease in weight is due to the expulsion of the water of hydration and calculate the per cent water of hydration by weight in bluestone.

3. From the chemical formula for bluestone, calculate the percent by weight of water of hydration in bluestone and compare the value obtained experimentally with this.

4. Calculate the percent error as follows:

$$\frac{\text{difference in percentage between (2) and (3)}}{\text{correct percentage as calculated}} \times 100\%$$

Conclusion:

1. What law of chemistry is illustrated?
2. Write the equation to represent the reaction taking place.
3. Enumerate any sources of error.
4. What conclusions can you draw from this experiment?

EXPERIMENT 4.

PERCENTAGE OF OXYGEN IN POTASSIUM CHLORATE

Object:

Large Pyrex test tube or small Erlenmeyer flask, balance, test-tube holder, Bunsen burner. (For flask a ring stand and wire gauze is required.)

Procedure:

(1) Weigh the empty test tube or Erlenmeyer flask.

(2) Place about two grams of dry potassium chlorate into the flask and weigh.

(3) Heat the potassium chlorate gently so that the oxygen is given off slowly. Guard against spattering.

(4) Continue heating until the bubbling ceases, then heat strongly for 5 minutes.

(5) Cool and weigh.

(6) Heat again for 5 minutes, cool and weigh.

(7) Repeat step 6 until a constant weight is attained.

Observations:

(1) Wt. of flask =

(2) Wt. of flask and potassium chlorate =

(3) Wt. of potassium chlorate =

(4) Wt. after first heating =

(5) Wt. after second heating =

(6) Wt. of oxygen =

(7) Per cent oxygen =
(the loss in weight is equal to the weight of oxygen).

(8) From the chemical formula of potassium chlorate calculate the percent by weight of oxygen present and compare the experimental value with this.

(9) Calculate the percent error.

Conclusion:

(1) Write the equation representing the reaction taking place.

(2) What law of chemistry is illustrated?

(3) State any possible sources of error.

EXPERIMENT 5

ACTIVITY OF METALS
(Teacher Demonstration)

Object:

(1) To determine the activity of various metals by displacement of hydrogen from water.

(2) To show that a base is formed when a metal displaces hydrogen from water.

Apparatus and Materials:

Potassium, sodium, calcium, copper, beakers, litmus solution.

Procedure:

(1) Fill each of 4 beakers about ½ full of water.

(2) Put a few drops of red litmus solution into each beaker.

(3) Drop a piece of potassium metal, no larger than the size of a pea, into the first beaker.

(4) Repeat (3), dropping a piece of sodium into the second beaker, calcium into the third and copper into the fourth.

Observation:

(1) How are the metals, sodium and potassium, stored? Why?

(2) What precautions were taken in cutting and handling potassium and sodium? Why?

(3) How is calcium stored?

(4) Why was copper not handled with as much care as sodium and potassium?

(5) Describe the reactions that took place in each of the beakers. Which reaction was most vigorous? Least, if any?

(6) Describe the color changes of the water solution.

(7) Touch the liquid and rub it between your fingers. does it feel? Wash your fingers with water after touching the liquid. Why?

Conclusion:

(1) Which metal is most active?

(2) Which metal is least active?

(3) Arrange the metals in order of activity.

(4) Which metal would you expect to find below hydrogen in the activity series?

(5) Name the resulting solutions in each beaker.

(6) Write equations for the reactions.

Note: To test that the gas given off is hydrogen, wrap the metal in perforated tin foil before dropping into water and collect the gas in a test tube by holding the test tube directly over the foil. Test with burning splint.

EXPERIMENT 6

THE ACTIVITY OF METALS

Object:

To determine the activity of various metals by displacement.

Apparatus and Materials:

Test tubes, 50 ml. beaker, iron nail, copper sulfate crystals, magnesium ribbon, zinc, splint, hydrochloric acid, sulfuric acid, Bunsen burner.

Procedure:

Part A.

(1) Place about one inch of magnesium ribbon in a Pyrex test tube and add about 10 ml. of dilute hydrochloric acid.

(2) Quickly place another test tube in an inverted position above this test tube and collect the gas.

(3) Test the gas with a burning splint. (Be careful that the test tube is directed away from yourself and not directed toward any other person.)

(4) Repeat steps (1) to (3) using a small piece of zinc.

Observation:

(1) Describe the method of collection of this gas. Why is it possible to collect it in this manner?

(2) Describe the physical properties of the gas (color, odor, etc.) and the results of introducing a burning splint.

Conclusion:

(1) What displacement took place and why?
(2) Identify the gas replaced.
(3) Write equations for the two reactions.

Procedure:

Part B.

(1) Dissolve three or four copper sulfate crystals in 25 ml. of water in the beaker.

(2) Note the physical properties of the solution and of a clean iron nail.

(3) Immerse the nail in the solution for five minutes.

(4) Examine the nail and note any changes.

Observation:

(1) Describe the appearance of the nail before and after being immersed in the solution.

Conclusion:

(1) Discuss what took place in terms of chemical activity.
(2) Write the equation representing the reaction.

Procedure:

Part C.

(1) Dissolve three or four copper sulfate crystals in 25 ml. of water.

(2) Add small pieces of zinc to the solution and stir vigorously until the color of the solution has almost disappeared.

(3) Allow the powder to settle; decant the supernatant liquid and discard it.

(4) Dry the powder by heating gently and note its properties.

Observation:

(1) Describe the changes taking place.

(2) Note the properties of the powder that remains.

Conclusion:

(1) Identify the powder which remains.

(2) Discuss what occurred in terms of chemical activity.

(3) Write the equation representing the reaction.

Arrange the elements encountered in the laboratory proce-

dure according to their activity and discuss the reasons for so arranging them.

EXPERIMENT 7.

Object:

To test the action of indicators in acids and bases.

Apparatus and Materials:

Test tubes, water, blue and pink litmus paper, methyl orange, methyl violet, phenolphthalein, hydrochloric, sulfuric and nitric acid, sodium hydroxide and ammonium hydroxide.

Procedure:

Part A.

(1) Place three clean test tubes in a test-tube rack and number them.

(2) Into the first test tube place 5 ml. of dilute hydrochloric acid, into the second 5 ml. of dilute sulfuric acid and into the third 5 ml. of dilute nitric acid.

(3) Into each test tube dip a piece of blue litmus paper.

(4) Remove the paper and record the results in a table similar to that on page 98 of the text.

(5) Into each of the test tubes add 3 or 4 drops of methyl orange. Record the results.

(6) Wash out the test tubes, add water and acid as before. Add three or four drops of methyl violet solution to each. Record the results.

(7) Repeat the procedure using phenolphthalein.

Procedure:

Part B.

(1) Into two of the clean test tubes place 10 ml. of water.

(2) Add 5 ml. of dilute sodium hydroxide to the first test tube and 5 ml. of dilute ammonium hydroxide to the second.

(3) Into the test tube dip a piece of red litmus paper.

(4) Remove the paper and record the results in tabular form.

(5) Add three or four drops of methyl orange to each test tube and record the results.

(6) Into two clean test tubes place the water and base as before.

(7) Follow the same procedure using methyl violet and

phenolphthalein.

(8) Record your results in the table.

Conclusion:

(1) From the experiment what conclusion would you make concerning the color of litmus in the presence of an acid? a base?

(2) What color is methyl orange in an acid solution? a basic solution?

(3) What colors are methyl violet and phenolphthalein in an acid solution? a basic solution?

EXPERIMENT 8.

NEUTRALIZATION (QUANTITATIVE)

Object:

To show the neutralization of an acid by a base.

Apparatus and Materials:

0.20N hydrochloric acid, sodium hydroxide, burette, Erlenmeyer flask, graduated cylinder, phenolphthalein.

Procedure:

(1) Clean a burette thoroughly and mount in a stand. Also clean an Erlenmeyer flask.

(2) Into the Erlenmeyer flask place 25 ml. of 0.20N hydrochloric acid and add a few drops of phenolphthalein.

(3) Fill the burette with the sodium hydroxide solution and take the initial reading.

(4) Slowly add the sodium hydroxide to the acid solution, agitating the acid by rotating the flask.

(5) As it appears that the end point is approaching (the color fades less quickly), add the sodium hydroxide dropwise.

(6) Continue adding sodium hydroxide until the first permanent pale pink color appears.

(7) Take the volume reading at the end point. (The burette should be washed immediately as sodium hydroxide tends to "freeze" the stopcock.)

Observations:

Record the observations as follows:

(1) volume of acid -

(2) normality of acid -

(3) initial volume reading -

(4) final volume reading -

(5) volume of base -

Use the relationship:

$$V_A N_A = V_B N_B$$

to calculate the normality of the base used.
Normality is defined as the number of gram-equivalents of solute per liter of solution.

Conclusion:

Write the equation for the reaction.

EXPERIMENT 9.

HYDROGEN CHLORIDE GAS AND HYDROCHLORIC ACID

(Teacher Demonstration)

Preparation and Collection of Hydrogen Chloride Gas:

Prepare hydrogen chloride gas by the reaction of concentrated sulfuric acid on sodium chloride.

Diagram: (See page 94)

Equation: . . . .

Why did you collect hydrogen chloride by the upward displacement of air?

(1) . . . .

(2) . . . .

Note some properties of the gas.

(1) Color . . . . (2) Odor. . . . .

(3) Solubility in water . . . .

Bubble some of the gas through water. Test with litmus.

Result . . . .

Name the solution . . . .

Test for hydrogen chloride gas.

Remove the glass stopper from a bottle containing hydrochloric acid and hold it near the glass stopper from a bottle containing ammonium hydroxide.

Observation . . . .

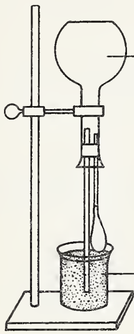
Equation . . . .

Explanation . . . .

Hydrogen Chloride fountain.

Collect some hydrogen chloride gas in a dry florence flask. Fit flask with a two-hole stopper. In one hole insert a medicine dropper filled with water. Fit a glass tube in the other hole and invert so that one end of glass tube is placed in a beaker of water to which a few drops of blue litmus solution were added.

Diagram:



Use medicine dropper to force a few drops of water into flask.

Observation . . . .

Conclusion . . . .

EXPERIMENT 10

TESTS FOR CHLORIDES

Object:

To prepare a chloride and to determine some of its properties.

Apparatus and Materials:

Cupric chloride, sodium chloride, silver nitrate, iron nails, test tubes, nitric acid, ammonium hydroxide, cupric oxide, hydrochloric acid.

Procedure:

Part A.

(1) In this experiment, a known solution is used for comparison. To set up the known solution, one gram of

anhydrous cupric chloride is dissolved in 15 ml. of water in a test tube. Mark this as test tube A.

(2) Take 0.3 gm. of cupric oxide and place it in a test tube the same size as test tube A.

(3) Add dilute hydrochloric acid slowly to the cupric oxide stirring until it is completely dissolved.

(4) Add water until the level of the second test tube is the same level as test tube A. Mark this as test tube B.

(5) Compare the color and appearance of the two test tubes.

(6) Immerse two clean iron nails, one in each test tube for a few minutes. Remove and note the results.

(7) Add silver nitrate to each of the solutions until precipitation is complete.

(8) Decant the supernatant liquid in each; add 10 ml. of distilled water to the precipitate in each to wash it. Decant the water.

(9) Divide the precipitate of test tube A into three parts:

 (a) To the first part add dilute ammonium hydroxide. Agitate and note the results.

 (b) To the second part add dilute nitric acid. Note the results.

 (c) Place the third portion in direct sunlight for a few minutes and note any changes.

(10) Repeat steps 9 a, b, and c for the precipitate in test tube B, first dividing it into three equal parts.

(11) Step 9 may be repeated with a sodium chloride solution to which silver nitrate has been added.

Observation:

Record all the observations in order.

Conclusion:

(1) Write equations to represent the reactions taking place in Steps 2, 6, 7, and 9 (a).

(2) In carrying out steps 9 (a), (b) and (c) with other chlorides, chemists have found that the results are characteristic of all chlorides. Hence this process is accepted as a test for chlorides. State the properties of silver chloride by which it is recognized.

SECTION III

UNIT 1 - A SHORT HISTORY OF SCIENCE

CHAPTER 1. OUR DEBT TO THE PAST

Page 10

Part A

1. The word "science" is derived from the Latin verb <u>scio</u> meaning "to know".

2. Progress in science was slow during the early period of its study because

    (a) communication was poor.

    (b) natural phenomena were put on a mystical basis.

    (c) often theories were not based on observation.

3. A false theory was a sign of progress, in that it encouraged other scientists to further their studies and make corrections.

4. The Babylonians were familiar with astronomy, to the extent that they could foretell the occurrence of a solar eclipse.

5. Some of the equipment used by alchemists is similar to present-day equipment. e.g. mortar and pestle, retort, Erlenmeyer flask, florence flask. An oven provided heat.

6. Paracelsus was one of the first to denounce the practice of alchemy. He noted a relationship between cause and effect, made a study of certain compounds and attempted to establish their use in medicine. His methods are still used today and represented a break from alchemy.

7. (a) The facts concerning the problem are collected by observation and experimentation.

    (b) A reasonable explanation of the observed reaction or phenomenon is formulated.

    (c) Results of related reactions are prophesied from the general theory.

    (d) The general theory is tested, modified and, if necessary, retested.

    Steps (a) and (b) employ inductive reasoning, whereas steps (c) and (d) employ deductive reasoning.

8. In inductive reasoning, known facts pertinent to the problem are collected and a theory formulated from them. It was first advocated by Francis Bacon.

9. In inductive reasoning one goes from specific facts to a general theory. In deductive reasoning, a general theory is used to interpret specific facts.

$$\text{specific} \rightleftarrows \text{general}$$

10. R. Boyle defined an element as a substance which could not decomposed into simpler substances by any means available at the time. (1661).

11. The scientist is accepted today because of his contributions to the standard of living and the general welfare of humanity, as exemplified by his contributions to medicine, industry, communication and the comforts of everyday living.

Part B

12. Aluminum and iron are rarely found in the free state and are difficult to refine. The ancients did not know how to refine them.

13. Alchemists did not base their theories on observation, but rather on a supernatural basis.

14. Alchemists believed that "base" metals such as lead could be changed into gold by means of a mysterious substance called the "philosopher's stone".
Modern scientists studied the nature of elements carefully and employed procedures indicated by these studies.

15. People were superstitious and the alchemists claimed to have mystical powers. People desired the things for which they were searching, e.g., eternal life and riches.

16. Electricity, modes of transportation and communication, the internal-combustion engine.

17. Electrical appliances, clean methods of heating, telephone, radio, television, car, airplane.

18. Modern machines shortened our hours of work. Electricity provides a good source of light and heat.

    Improvements in medicine.
    Transportation methods were improved.
    Provided more luxuries.

19. Aristotle's writings, after being accepted by the church, were considered the sole authority for hundreds of years.
Opposition encountered to the fluoridation of water.

20. A young Canadian discovered that inert elements are not really inert.

21. This method can be applied to projects involving the studies of the political and economic structure of a community.

UNIT 2 - MATTER

CHAPTER 2. MATTER: ITS CLASSIFICATION

Page 26

Part A

1. Matter is that which has weight and occupies space. Strictly speaking, it is an undefined term used by scientists. Invisible, very small particles of matter, e.g. gases are difficult to recognize with the unaided eye.

2. Solid, liquid, gas.

3. A solid has definite shape and volume whereas a liquid has a definite volume but no definite shape. Gases have neither a definite shape nor volume.

4. In a "pure substance" all particles are alike. Elements and compounds are pure substances.

5. An element is a substance which cannot be decomposed into simpler substances by chemical means.

6. A compound is the product of the chemical union of two or more elements.

7. In a mixture, individual particles may differ; mechanical mixtures, solutions, alloys.

8. Physical properties deal with the form and appearance of the element.

9. Chemical properties deal with the ability or lack or ability of the element to react with other elements.

10. Solid, dark gray, density 7.75 gm/cc, malleable, somewhat ductile, hard, melting point 1530^{o}C.

11. State color, hardness, density, taste, odor, malleability, ductility, solubility, melting point, boiling point.

12. (a) physical - solubility is a physical property.
    (b) physical - change of state.
    (c) physical - change of state.
    (d) chemical - new substance formed, combines with oxygen.
    (e) physical - change of state.
    (f) chemical - slight change in composition.
    (g) chemical - new substance formed.
    (h) physical - solution lowers the melting point.

13. Exothermic reactions liberate heat.

14. Dilution of sulfuric acid.

15. Endothermic reactions absorb heat.

16. - $NH_4C_2H_3O_2$ - negative heat of solution.

17. Energy is the ability to do work.

    (a) kinetic energy - energy due to motion.
    (b) potential energy - energy due to position.
    (c) mechanical energy - energy of matter moving against matter
    (d) electrical energy - energy possessed by flowing electrons.

18. Meter; centimeter and millimeter.
    Gram, liter. At 4^{o} C. and 760 mm. pressure, 1 cubic centimeter (or milliliter) of water weighs 1 gram.

19. It is a decimal system, i.e. Related units are all multiples of ten.

20. $220 \text{ lb.} = 220 \text{ lb.} \times \frac{1 \text{ kg.}}{2.2 \text{ lb}} = 100 \text{ kg.}$

or

$1 \text{ kg.} = 2.2 \text{ lb.}$

$\frac{1}{2.2} \text{ kg.} = 1 \text{ lb.}$

$\frac{220}{2.2} \text{ kg.} = 220 \text{ lb.}$

$200 \text{ lb.} = 100 \text{ kg.}$

21. 1 m. = 39.37 in.

1 m. = 100 cm.

100 cm. = 39.37 in.

$100 \text{ cm.} = \frac{39.37}{12} \text{ ft.} = 3.28 \text{ ft.}$

22. 1 kg. = 1000 gm.

12.25 kg. = 12,250 gm.

23. 1 litre = 1000 ml.

11.55 l. = 11,550 ml.

24. 1000 cc. = 1 l.

$6,475,35 \text{ cc.} = \frac{6,475.35 \text{ l.}}{1000} = 6.47535 \text{ l. or } 6.48 \text{ l.}$

25. 1 l. = 0.88 qt.

$\frac{1}{0.88} \text{ l.} = 1 \text{ qt.}$

$\frac{44}{0.88} \text{ l.} = 44 \text{ qt.}$

50 l. = 44 qt.

Part B

26. Solids are heated until they melt, mixed together, and allowed to cool. Steel is formed by heating a mixture of iron and carbon.

27. Two liquids are said to be miscible if they are soluble in each other, e.g. water and alcohol.

28. Physical: there is no change in the composition of the nail.

29. (a) Wax melts; wood does not melt.
(b) Density, color, elasticity.
(c) Color, density, chemical reactions.
(d) Density, chemical reactions.
(e) Odor, gasoline burns, water doesn't.

30. Add water - salt dissolves in water. Evaporate water, thus recovering the salt.

31. Copper: reddish brown, malleable, lustrous.
Sulphur: yellow, finely divided, flammable.

32. Steel is magnetic whereas lead is not.

33. On the basis of density. If the container is agitated the heavier lead shot would remain at the bottom.

34. Center of the earth; on the sun.

35. Burning of wood, burning of coal; rusting of iron; cooking of food, decaying of organic matter.

36. To minimize evaporation; to retain any volatile materials in case any are formed.

37. No. This is an example of sublimation, i.e. the change from the solid state to the gaseous state without passing through the intermediate liquid state.

38. Weigh one, then weigh them all.

CHAPTER 3. MATTER: ITS COMPOSITION

Page 37.

Part A

1. Today's definition includes the phrase "by chemical means".

2. This phrase is included because it has been found that elements can be decomposed into simpler particles by other means.

3. 17th Century.

4. Radioactive decay is not a chemical change. The change involves changes in the nucleus.

5. Mercury. Oxygen, or nitrogen, or chlorine or fluorine; sulphur; bromine.

6. Methods of refining aluminum in quantity were unknown.

7. The percentage composition of a compound is fixed whereas that of a mixture can be altered. A compound does not have the properties of its constituents whereas a mixture does.

8. The properties of a compound differ from those of the elements making up the compound.

9. The composition of a compound by weight is fixed.

10. All samples of any compound contain the same elements in the same fixed proportion by weight.

11. An atom consists of a central nucleus consisting of protons and neutrons. Surrounding this are electrons. The diameter of the nucleus is very small conpared to that of the whole atom, but most of the mass of the atom is concentrated in the nucleus.

12. One atom of the substance; one atomic weight of the substance.

13. It was chosen as it is the lightest of all elements.

14. Oxygen is more abundant and its compounds are more readily formed than those of hydrogen. Oxygen taken as 16.000 eliminates fractions in many of the other elements.

15. Berzelius.

16. Communication is facilitated.

Part B

17. A molecule is the smallest part of an element or compound that possesses the properties of that element or compound and can exist by itself. O_2 ; H_2O

18. K; O.

19. No; water molecules consist of two hydrogen atoms and one oxygen atom each.

20. Dalton believed atoms were the smallest particles of matter; that atoms were indivisible and that all the atoms of a single element were alike with respect to shape, size, and weight. Today we know atoms are not the smallest particles that exist; that they can be subdivided, and that atoms of a single element are not all alike, but exist as isotopes.

21. Gram-molecular weights are strictly ratios.
An oxygen molecule is 16 times the weight of a hydrogen molecule. Later on, it will be shown that one gram-molecular weight represents the weight of 22.4 liters of a gas or vapor at S.T.P.

CHAPTER 4. MATTER: ITS STRUCTURE

Page 47

1. The beam of light travelling from cathode to anode could be deflected by a magnet.

2. Sir J.J. Thomson.

3. This ray consisted of tiny negatively-charged particles of matter called electrons.

4. An electron is a small negatively-charged particle of matter.

5. By bombarding a metal foil with alpha particles and found that some were deflected at fairly large angles, indicating a dense positive nucleus.

6. There existed a discrepancy in the weight of all atoms except hydrogen with the weight of the protons (and electrons) present.

7. A neutron is a small particle of matter which is not electrically charged.

8. Sir James Chadwick.

9. The total weight of an atom is composed of the weight of the protons, neutrons and electrons.

10. Mendeleef.

11. Argon and potassium; tellurium and iodine; cobalt and nickel.

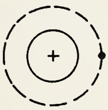
12. The atomic number indicates the number of protons present in the nucleus of the atom.

13. The number of neutrons is the difference between the atomic weight and atomic number of the element.

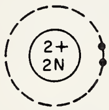
14. 2.

15. 8.

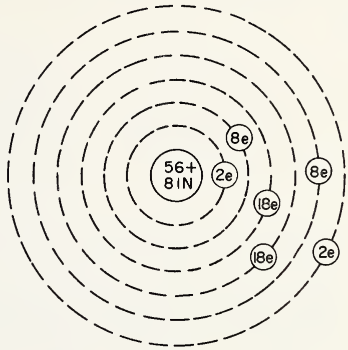
16.



H°

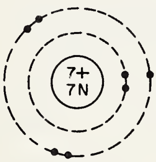


He

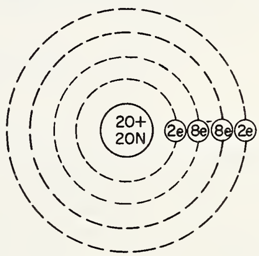


Ba°

C°



N°



Ca°

17. It is not according to scale; it is not three dimensional; it gives us the impression that electrons are found in circular orbits.

18. Millikan determined the mass of the electron to be 1/1837 the mass of a proton.

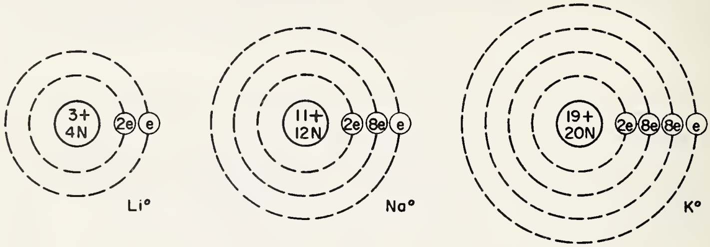
Part B

19. Fluorine will show the greatest attraction for the electron because the atom is small and the positive nucleus is not shielded to the same extent by planetary electrons as are those of the others.

20. Potassium will lose its electron most readily as the distance between it and the nucleus is greater than in the other two cases. Also the shielding effect of the other electrons tends to weaken the attraction between the nucleus and the outermost electron.

21. The chemical properties of elements depend upon their atomic numbers. Elements with similar properties appear at regular intervals.

22.



Each has one electron in the outer orbit.

23. Chlorine is an active non-metal as it lacks one electron to complete its outer shell. Argon is inert as its outer shell is "stable" and potassium, with one electron in the outer shell is likely to lose this electron to attain a stable octet structure.

UNIT 3 - THE LANGUAGE OF CHEMISTRY

CHAPTER 5. SYMBOLS AND FORMULAE

Page 65

1. A symbol also represents one atom of an element and one atomic weight of an element.
2. A formula represents the name of the molecule, one molecule of the substance and one molecular weight of the substance.
3. Arrhenius.
4. The theory of ionization.
5. Sodium has "lost" an electron, thus leaving it with an excess positive charge; chlorine has "gained" an electron, thus acquiring a negative charge.
6. The charges are produced by a transfer of electrons.
7. The calcium atom has two electrons in the outer orbit, which it loses in forming a compound.
8. Oxygen has six electrons in the outer shell and needs to gain two in order to attain a stable octet structure.
9. A "stable octet" consists of eight electrons in the outer orbit and is particularly resistant to further loss or gain of electrons.

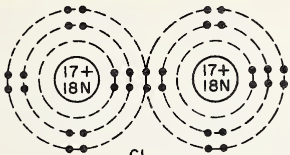
10. Neutrality refers to a lack of either excess positive or negative charge, while stability refers to the resistance of an element to gain or lose electrons. Atoms are neutral, while ions are stable.

11. A non-metal will achieve stability by the gain of electrons.

12. A metal will achieve stability by the loss of electrons.

13. The sharing of electrons.

14.



Cl_2

15. Metals usually have fewer than half the required number of electrons for stability in the outer orbit so they will lose these electrons, making the previous orbit the outer shell. Non-metals usually have more than half the required number of electrons for stability in the outer orbit so electrons will be gained to complete the orbit.

16. The number of electrons gained, lost or shared is called the valence. Hydrogen can lose or share one electron so its valence is +1. Chlorine can gain or share one electron so its valence is -1.

17. Helium already has a complete outer shell, so it will not lose or gain or share electrons.

18. The valence will indicate the number of atoms of the element necessary in order to form a neutral compound.

19. The ring structure of copper may be either 2, 8, 18, 1 or 2, 8, 17, 2.

20. The suffix -ous indicates the lower valence (+2 for iron) while the suffix -ic indicates the higher valence (+3 for iron).

21. The suffix -ide.

22. (a) $CaCl_2$ - calcium chloride.

    (b) NaBr - sodium bromide.

    (c) AlN - aluminum nitride.

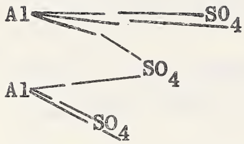
    (d) K_2S - potassium sulfide.

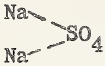
23. A radical is a stable group of atoms possessing a definite valence that acts as a single unit in most chemical reactions.

24. (a) PO_4^{---} (b) SO_4^{--} (c) CO_3^{--} (d) HCO_3^{-}

    (e) HSO_4^{-} (f) ClO_3^{-} (g) NH_4^{+}

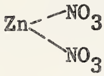
Part B

25. 

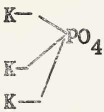
26. (a) 

Na_2SO_4 - sodium sulfate

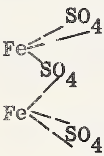
(b) Cu = CO_3 $CuCO_3$ - cuprous carbonate

(c) 

$Zn(NO_3)_2$ - zinc nitrate

(d) 

K_3PO_4 - potassium phosphate

(e) 

$Fe_2(SO_4)_3$ - ferric sulfate

(f) K - - ClO_3 $KClO_3$ - potassium chlorate

27. Recall a zinc compound whose formula and valence of the anion you remember. From that you can work out the valence of zinc.

28. Sugar consists of the elements carbon, hydrogen, and oxygen. There are 12 carbon atoms, 22 hydrogen atoms and 11 oxygen atoms present in the molecule. Its gram molecular weight is 342 gm.

29. (a) a transfer of a single electron, resulting in electrostatic attraction.
(b) a sharing of one pair of electrons.

30. N - one atom of nitrogen, atom wt. 14.
N_2 - one molecule of nitrogen, molecular wt. 28.
$2N_2$ - two molecules of nitrogen, total molecular weight of 56.

31.

Copper	Sulfur
Reddish-brown	Yellow
luster	not lustrous
malleable	not malleable
ductile	not ductile

Copper	Sulfur
non-flammable	inflammable
relatively high melting point	low melting point
high density	finely divided
good conductor of electricity and heat	low density
	poor conductor of electricity and heat

CHAPTER 6. CHEMICAL EQUATIONS

Page 77

1. An equation is a description of a reaction that actually takes place.
2. $2HgO \longrightarrow 2Hg + O_2$
3. The total number of atoms of an element present in the elements or compounds entering the reaction must equal the total number of atoms of an element present in the products formed.
4. Subscripts indicating the number of atoms of an element present cannot be changed at will as the composition of a molecule is constant.
5. A molecule of mercuric oxide consists of one atom of mercury and one atom of oxygen. Any change in the subscripts would violate the law of constant proportions.
6. The formulae of the reactants are written on the left side and the molecular formulae of the products formed on the right side. The total number of atoms of each element must be the same on both sides of the equation.

$$2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + 2H_2O$$

7. A catalyst is a substance which changes the rate of a chemical reaction without itself undergoing a chemical change.
8. Yes. MnO_2 is a catalyst in the preparation of oxygen but a reactant in the preparation of chlorine.
9. (a) direct synthesis - burning hydrogen in air.
   (b) simple replacement - magnesium ribbon is placed in hydrochloric acid.
   (c) double decomposition - sulfuric acid is added to limestone
   (d) simple decomposition - electrolysis of acidulated water.
10. Equations may be used to determine the weights (or volumes) of products produced from definite weights (or volumes) of reactants. Two gram molecular weights (or some proportion thereof) of potassium chlorate will produce two gram molecular weights (or some proportion thereof) of potassium chloride and three molecular weights (or some proportion thereof) of oxygen gas.

Part B

11. (a) $2K + H_2SO_4 \longrightarrow K_2SO_4 + H_2$ (simple replacement)
    (b) $CaCO_3 \longrightarrow CaO + CO_2$ (simple decomposition)
    (c) $CuO + H_2 \longrightarrow Cu + H_2O$ (simple replacement)
    (d) $Cu + Cl_2 \longrightarrow CuCl_2$ (direct synthesis)
    (e) $2H_2 + O_2 \longrightarrow 2H_2O$ (direct synthesis)
    (f) $2Na + 2H_2O \longrightarrow 2NaOH + H_2$ (simple replacement)

12. A molecule of oxygen consists of two oxygen atoms while only one oxygen atom is present in mercuric oxide.

13. It expresses the quantitative relationship that the total weight of the products is equal to the total weight of the reactants.

14. Catalyst required, whether the substances are dry or in solution, concentrations, conditions under which the reaction takes place, e.g. the temperature required for the reaction to take place.

15. The reactants and products and their formulae must be known.

16.
$$\underset{56\text{ gm.}}{\overset{14\text{ gm.}}{Fe}} + \underset{32\text{ gm.}}{\overset{8\text{ gm.}}{S}} \longrightarrow FeS$$

17.
$$\underset{56\text{ gm.}}{\overset{11.2\text{ gm.}}{Fe}} + \underset{32\text{ gm.}}{\overset{x}{S}} \longrightarrow \underset{88\text{ gm.}}{\overset{y}{FeS}}$$

(a) $\frac{11.2\text{ gm.}}{56\text{ gm.}} = \frac{x}{32\text{ gm.}}$

$x = \frac{11.2\text{ gm.} \times 32\text{ gm.}}{56\text{ gm.}}$

$= 6.4$ gm.

6.4 gm. of sulfur are required.

(b) (i) There are 11.2 gm. + 6.4 gm. = 17.6 gm. FeS formed.
(ii) $\frac{11.2\text{ gm.}}{56\text{ gm.}} = \frac{y}{88\text{ gm.}}$

$y = \frac{11.2\text{ gm.} \times 88\text{ gm.}}{56\text{ gm.}} = 17.6$ gm.

17.6 gm. of FeS are formed.

18.
$$\underset{245\text{ gm.}}{\overset{x\text{ gm.}}{2\,KClO_3}} \longrightarrow 2\,KCL + \underset{96\text{ gm.}}{\overset{8\text{ gm.}}{3\,O_2}}$$

(continued)

18. (cont'd).

$$\frac{x \text{ gm.}}{245 \text{ gm.}} = \frac{8 \text{ gm.}}{96 \text{ gm.}}$$

$$x = \frac{\cancel{8} \times 245}{\underset{12}{\cancel{96}}} = 20.4$$

∴ 20.4 gm. of potassium chlorate are required.

19.

$$\begin{array}{ccccc} x \text{ gm.} & & & & 8 \text{ gm.} \\ 2HgO & \longrightarrow & 2Hg & + & O_2 \\ 432 \text{ gm.} & & & & 32 \text{ gm.} \end{array}$$

$$\frac{x \text{ gm.}}{432 \text{ gm.}} = \frac{8 \text{ gm.}}{32 \text{ gm.}}$$

$$x = \frac{8 \times 432}{32} = 108.$$

∴ 108 gm. of mercuric oxide are required.

20.

$$\begin{array}{ccccccc} 80 \text{ gm.} & & 2 \text{ gm.} & & 64 \text{ gm.} & & 18 \text{ gm.} \\ CuO & + & H_2 & \longrightarrow & Cu & + & H_2O \\ 80 \text{ gm.} & & 2 \text{ gm.} & & 64 \text{ gm.} & & 18 \text{ gm.} \end{array}$$

(a) 2 gm. of hydrogen would be required
(b) 18 gm. of water would be formed
(c) 64 gm. of copper would be produced

21. H_2SO_4 - molecular weight is 2 + 32 + 64 = 98

$$\% H = \frac{2}{98} \times 100\% = 2.04\%$$

$$\% S = \frac{32}{98} \times 100\% = 32.65\%$$

$$\% O = \frac{64}{98} \times 100\% = 65.31\%$$

22. $BaCl_2 \cdot 2H_2O$ molecular wt is 244

$$\% H_2O = \frac{36}{244} \times 100\% = 14.75\%$$

23.

$$\begin{array}{ccccc} x \text{ gm.} & & & & 110 \text{ gm.} \\ CaCO_3 & + & H_2SO_4 & \longrightarrow & CaSO_4 + H_2O + CO_2 \\ 100 \text{ gm.} & & & & 44 \text{ gm.} \end{array}$$

$$\frac{x \text{ gm.}}{100 \text{ gm.}} = \frac{110 \text{ gm.}}{44 \text{ gm.}}$$

$$x = \frac{110 \times 100}{44} = 250$$

∴ 250 gm. of $CaCO_3$ would be required.

24. In 392 gm of 75% sulfuric acid there are 0.75 x 392 gm = 294 gm pure sulfuric acid.

$$\underset{}{CaCO_3} + \underset{98\text{ gm}}{\overset{294\text{ gm}}{H_2SO_4}} \longrightarrow CaSO_4 + H_2O + \underset{44\text{ gm}}{\overset{x\text{ gm}}{CO_2}}$$

$$\frac{294\text{ gm}}{98\text{ gm}} = \frac{x\text{ gm}}{44\text{ gm}}$$

$$x = \frac{44 \times 294}{98} = 132$$

.˙. 132 gm of carbon dioxide would be produced.

25. (a)

$$\underset{65\text{ gm}}{\overset{x\text{ gm}}{Zn}} + \underset{71\text{ gm}}{\overset{182.5\text{ gm}}{2HCl}} \longrightarrow ZnCl_2 + H_2$$

$$\frac{x\text{ gm}}{65\text{ gm}} = \frac{182.5\text{ gm}}{\cancel{36.5}\text{ gm}}$$
71

$$x = \frac{182.5 \times 65}{36.571} = 162.5$$

.˙. 162.5 gm of zinc are required.

(b)

$$\underset{56\text{ gm}}{\overset{x\text{ gm}}{Fe}} + \underset{71\text{ gm}}{\overset{182.5\text{ gm}}{2HCl}} \longrightarrow FeCl_2 + H_2$$

$$\frac{x\text{ gm}}{56\text{ gm}} = \frac{182.5\text{ gm}}{71\text{ gm}}$$

$$x = \frac{182.5 \times 56}{71} = 140$$

.˙. 140 gm of iron are required.

(c)

$$\underset{54\text{ gm}}{\overset{x\text{ gm}}{2\ Al}} + \underset{219\text{ gm}}{\overset{182.5\text{ gm}}{6HCl}} \longrightarrow 2AlCl_3 + 3H_2$$

$$\frac{x\text{ gm}}{54\text{ gm}} = \frac{182.5\text{ gm}}{219\text{ gm}}$$

$$x = \frac{182.5 \times 54}{219} = 45$$

.˙. 45 gm of aluminum are required.

UNIT 4 - SOME SUBSTANCES MAN HAS LEARNED TO USE

CHAPTER 7. OXYGEN

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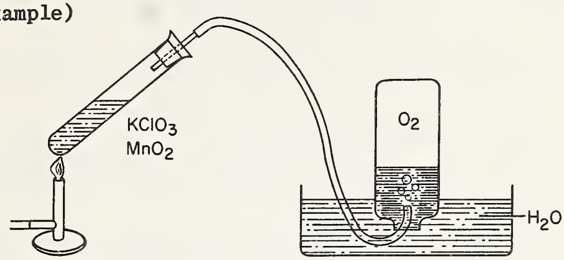
1. Oxygen (one-fifth by volume), nitrogen (four-fifths by volume), and small amounts (about 1%) of argon, Xenon, neon, helium, krypton, carbon dioxide and water vapor.
2. It is necessary for life.

3. Oxides.

4. A glass container filled with mercury and a little mercuric oxide was inverted in a glass dish containing mercury. Heat was provided by the sun's rays, concentrated by a lens.

5. (a) the modern system of nomenclature.
   (b) the introduction of quantitative procedures to chemistry.

6. (for example)



7. A catalyst is a substance which alters the rate of a chemical reaction without itself undergoing a chemical change.

8. Only small quantities are required. Time is a major factor, especially if the laboratory periods are short.

9. Raw materials are used up in the process. The equipment can be used over and over again.

10. A rise is temperature accompanies the compression of a gas. As a gas expands, its temperature is lowered. A sufficient lowering in temperature causes gases to liquify. Different liquids have different boiling points and can be separated on the basis of boiling point differences.

11. Colorless, odorless, slightly heavier than air (32 : 28.9), slightly soluble in water, liquifies at -183°C. and solidifies at -219°C.

12. Active, forms oxides with metals and with non-metals.

13. $2Hg + O_2 \longrightarrow 2HgO$

    $2Mg + O_2 \longrightarrow 2MgO$

    $4Al + 3O_2 \longrightarrow 2Al_2O_3$

    $2Zn + O_2 \longrightarrow 2ZnO$

    $S + O_2 \longrightarrow SO_2$

    $C + O_2 \longrightarrow CO_2$

14. This question is designed to encourage students to organize scientific information.

15. Liquid oxygen explosives.

16. The metal could be painted, galvanized or covered with tin.

17. $2H_2 + O_2 \longrightarrow 2H_2O$

In this example hydrogen is burned in oxygen. Oxygen supports the combustion of hydrogen. The fact that oxygen does not burn may de demonstrated by attempting to ignite oxygen as it is produced.

18. If wood is in the form of shavings, a greater surface area is exposed to the oxygen of the air. More heat is required to raise a greater mass to its kindling point.

19. Aquatic life uses the oxygen dissolved in water. As only small amounts dissolve in the laboratory the gas may be collected by downward displacement of water.

20. By "fanning", more oxygen from the surrounding air is forced to circulate among the coals.

21. As carbon unites with the oxygen of the air, the presence of greater quantities of oxygen would cause the combustion to be accelerated.

22. In rapid oxidation noticeable quantities of heat and light are produced, whereas this is not so apparent in slow oxidation.

23. More; iron unites with a definite weight of oxygen, rendering the compound formed heavier. Most of the oxides are gases and would escape to the atmosphere; thus reducing the weight of the product to be weighed. If all the gaseous products would be collected, then the weight would be more.

24.
$$\underset{48\text{ gm}}{\overset{10\text{ gm}}{2\,Mg}} + O_2 \longrightarrow \underset{80\text{ gm}}{\overset{X\text{ gm}}{2MgO}}$$

$$\frac{10\text{ gm}}{48\text{ gm}} = \frac{x\text{ gm}}{80\text{ gm}}$$

$$x\text{ gm} = \frac{10\ \cancel{\text{gm}}}{\underset{6}{\cancel{48}}\ \cancel{\text{gm}}} \times \overset{10}{\cancel{80}}\ \cancel{\text{gm}}$$

$$x = 16.67$$

$\therefore$ 16.67 gm of magnesium oxide would be produced.

Original magnesium weighed 10 gm.

The increase is $\frac{6.67}{10} \times 100\% = 66.7\%$ in weight over the original magnesium.

25.
$$\underset{432\text{ gm}}{\overset{x\text{ gm}}{2HgO}} \longrightarrow \underset{32\text{ gm}}{\overset{100\text{ gm}}{2Hg + O_2}}$$

$$\frac{x\text{ gm}}{432\text{ gm}} = \frac{100\text{ gm}}{32\text{ gm}}$$

$$x = \frac{100 \times 432}{32} = 1350\text{ gm}$$

(continued)

25. (Cont'd)

At \$5.00 per 100 gm., 1350 gm. would cost \$67.50.

$$\underset{245\ gm}{\overset{X\ gm}{2KClO_3}} \longrightarrow 2KCl + \underset{96\ gm}{\overset{100\ gm}{3O_2}}$$

$$\frac{x\ gm}{245\ gm} = \frac{100\ gm}{96\ gm}$$

$$x = \frac{100 \times 245}{96} = 255.2$$

At \$3.00 per kg. $\frac{255}{1000} \times \$3.00 = \$\ 0.77.$

255 gm would cost 77 cents.
The difference in cost would be \$66.73.

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26. $2C_8H_{18} \quad + \quad 25O_2 \longrightarrow 16CO_2 \qquad 18H_2O$

CHAPTER 8. HYDROCHLORIC ACID

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1. The word "acid" is derived from the Latin word <u>acidus</u> meaning "sour".
2. Citrus fruits, grapes, apples, milk, pickles (vinegar).
3. Hydrochloric acid; it is necessary for the proper digestion of protein.
4. Sour; affect indicators; react with most metals to produce hydrogen and a salt; neutralize bases.
5. Complex organic substances which react with acids (and bases) to produce color changes.
6. Blue litmus turns pink; methyl orange turns red.
7. Strong acids will ionize to a large degree; hydrochloric acid, sulfuric acid.
8. Weak acids will not ionize to any great extent; acetic acid, chloric acid.
9. A concentrated acid is one which is in solution with a small amount of water.
10. Dilute.
11. Hydrochloric acid ionizes to produces a high concentration of hydrogen ions and thus exhibits to a marked degree the properties common to all acids.
12. Hydrochloric acid is prepared by the action of concentrated sulfuric acid on sodium chloride to produce HCl gas. HCl gas dissolves in water to form HCl acid.

13. In the industrial preparation, a higher temperature is used to drive the reaction between sodium chloride and sulfuric acid to completion.

$2NaCl = H_2SO_4 \longrightarrow Na_2SO_4 + 2HCl$

14. The gas should not be inhaled.

15. Most acids may be prepared by the action of concentrated sulfuric acid on a salt of the desired acid.

16. A soluble oxide of a non-metal that will react with water to form an acid.

17. Carbon dioxide CO_2; sulfur trioxide SO_3.

$$CO_2 + H_2O \longrightarrow H_2CO_3$$

$$SO_3 + H_2O \longrightarrow H_2SO_4$$

18. (a) Sour taste
(b) Turns blue litmus pink; turns methyl orange red
(c) Reacts with many metals to produce hydrogen
(d) Neutralizes bases

19. The union of hydrogen and a non-metal or non-metallic radical which will ionize to form hydrogen ions.

20. "Pickling"; the cleaning of the base metal before plating; starting point for many chlorides; catalyst in some organic reactions.

21. Sulfuric acid.

22. It could be tested on the basis of conductivity.

23. (a) Hydrochloric acid, HCl
$HCl \rightleftarrows H^+ + Cl^-$

(b) Sulfuric acid, H_2SO_4
$H_2SO_4 \rightleftarrows 2H^+ + SO_4$

(c) Nitric acid, HNO_3
$HNO_3 \rightleftarrows H^+ + NO_3^-$

(d) Phosphoric acid H_3PO_4
$H_3PO_4 \rightleftarrows 3H^+ + PO_4^{---}$

(e) Chloric acid, $HClO_3$
$HClO_3 \rightleftarrows H^+ + ClO_3^-$

24. The moistened soil would be tested with litmus paper. If blue litmus turns pink, then the soil is acidic.

CHAPTER 9. SODIUM HYDROXIDE

Page 103

1. The action of metallic sodium with water.

$$2\ Na + 2HOH \longrightarrow 2NaOH + H_2$$

2. By means of an indicator.
3. The action of water on a basic anhydride.

$$Na_2O + HOH \longrightarrow 2NaOH$$

$$K_2O + HOH \longrightarrow 2KOH$$

$$CaO + HOH \longrightarrow Ca(OH)_2$$

4. Feel slippery; react with indicators to produce a color change; neutralize acids.
5. Sodium hydroxide ionizes to produce a high concentration of hydroxyl ions and thus exhibits to a marked degree the properties common to all bases.
6. The reaction of an acid and a base to produce a salt and water.
7. Every molecule consists of the characteristic part of the acid (H^+) and the characteristic part of a base (OH^-).
8. Litmus changes color at a point closer to a neutral solution than do the others.
9. It is produced by the electrolysis of brine. Sodium hydroxide is the primary product and hydrogen and chlorine are the by-products.
10. A physical process by which a substance absorbs sufficient moisture from the air to put itself in a concentrated solution.
11. A deliquescent substance should be stored in an air tight container.
12. As it is caustic it should be handled with forceps.
13. Sodium hydroxide is used as a starting point for the manufacture of other sodium compounds. It is also used as a drying agent.
14. Lye, caustic soda.
15. Sodium hydroxide is used in mercerizing thread and the manufacture of rayon.
16. Mercerizing involves the treatment of cotton fibers with sodium hydroxide and their drying under tension.
17. Sodium hydroxide is used in the manufacture of soap.
18. Sodium hydroxide is used to neutralize the sulfuric acid which was added to remove certain impurities.

$$2NaOH + H_2SO_4 \longrightarrow 2H_2O + Na_2SO_4$$

19. Wood pulp is treated with sodium hydroxide.

Part B

20. Sodium hydroxide is produced from sodium chloride which is inexpensive

21. Sodium is stored under kerosene in order to exclude any air, as it would react with the water vapor present in the air. Kerosene is a hydrocarbon. It contains no oxygen.

22. Sodium is a very active metal and forms compounds very readily, even with water vapor in the air.

23. Sodium chloride is inexpensive.

24. The end point would most likely be exceeded and a caustic burn result. An acid burn can be treated with sodium bicarbonate solution.

25. The solids could be dissolved in water and tested with an indicator.

26. No; the oxide must be soluble.

27. $2KOH + H_2CO_3 \longrightarrow K_2CO_3 + 2H_2O$

$2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + 2H_2O$

$3NaOH + H_3PO_4 \longrightarrow Na_3PO_4 + 3H_2O$

$3KOH + H_3PO_4 \longrightarrow 3H_2O + K_3PO_4$

$Ca(OH)_2 + 2HCl \longrightarrow CaCl_2 + 2H_2O$

$3Ca(OH)_2 + 2H_3PO_4 \longrightarrow Ca_3(PO_4)_2 + 6H_2O$

28.
$$\underset{46\ gm}{\overset{92\ gm}{2Na}} + 2H_2O \longrightarrow \underset{80\ gm}{\overset{x\ gm}{2NaOH}} + \underset{2\ gm}{\overset{y\ gm}{H_2}}$$

(a)
$$\frac{92\ gm}{46\ gm} = \frac{x\ gm}{80\ gm}$$
$$x = \frac{80 \times 92}{46} = 160$$

$\therefore$ 160 gm of sodium hydroxide would be formed.

(b)
$$\frac{96\ gm}{46\ gm} = \frac{y\ gm}{2\ gm}$$
$$y = \frac{2 \times 92}{46} = 4$$

$\therefore$ 4 gm of hydrogen would be produced.

29.
$$\underset{74\ gm}{\overset{x\ gm}{Ca(OH)_2}} + \underset{98\ gm}{\overset{147\ gm}{H_2SO_4}} \longrightarrow 2H_2O + CaSO_4$$

(continued)

$$\frac{x\ gm}{74\ gm} = \frac{147\ gm}{98\ gm}$$

$$x = \frac{\overset{3}{\cancel{147}} \times 74}{\underset{2}{\cancel{98}}} = 111\ gm$$

.'. 111 gm of calcium hydroxide would be required.

30. 5 gm sulfuric acid per gallon means that 2500 gm would be present in 500 gallons.

$$\underset{80\ gm}{\overset{x\ gm}{2NaOH}} + \underset{98\ gm}{H_2SO_4} \longrightarrow 2H_2O + NaSO_4$$

$$\frac{x\ gm}{80\ gm} = \frac{2500\ gm}{98\ gm}$$

$$x = \frac{80 \times 2500}{98} = 2040.8$$

.'. 2040.8 gm of caustic soda would be required.

CHAPTER 10 - SODIUM CHLORIDE

Page 110

1. Preparation and preserving of food, tanning leather, making pottery. A source of other sodium compounds; manufacture of drugs, dyes and fertilizers.
2. New Brunswick, Ontario, Northern Alberta and Saskatchewan.
3. Rock salt; chemically impure.
4. Halite may be blasted and brought out in blocks; a well may be drilled, water forced down to dissolve the salt and then be pumped out as brine.
5. Halite may be either blasted or dissolved (Ques. 4). Brine is treated to remove impurities and the water is boiled off in vacuum pans at a reduced pressure. The salt is then dried and graded.
6. Sodium chloride is necessary for the normal functioning of the blood and digestive juices.
7. Manufacture of sodium compounds, drugs, pottery glazes, synthetic rubber, dyes and fertilizers.
8. (a) Neutralization
   $HCl + NaOH \longrightarrow NaCl + H_2O$

   (b) Action of a metal and a non-metal.

   $Cu + S \longrightarrow CuS$

   (c) Action of a concentrated acid on the salt of a more volatile acid.

   $$H_2SO_4 + 2NaCl \longrightarrow Na_2SO_4 + 2HCl$$

(d) Action of a basic anhydride and acid anhydride.

$$CO_2 + CaO \longrightarrow CaCO_3$$

(e) Reaction of an acid with the oxide of a metal.

$$2HCl + CuO \longrightarrow CuCl_2 + H_2O$$

(f) The reatcion of two soluble salts to form a precipitate of another salt.

$$AgNO_3 + NaCl \longrightarrow NaNO_3 + AgCl$$

9. White, crystalline solid, odorless, soluble in water, reacts with sulfuric acid to form hydorgen chloride gas.

10. Cupric oxide and hydrochloric acid will react to produce cupric chloride and water. Hydrogen will replace the copper from the oxide.

11. (a) $Zn + 2HCl \longrightarrow ZnCl_2 + H_2$ Zinc Chloride.

(b) $2NaCl + H_2SO_4 \longrightarrow Na_2SO_4 + 2HCl$ Sodium Sulfate.

(c) $NaOH + HNO_3 \longrightarrow NaNO_3 + H_2O$ Sodium Nitrate.

(d) $3Ca(OH)_2 + 2H_3PO_4 \longrightarrow Ca_3(PO_4)_2 + 6H_2O$ Calcium phosphate.

12.

$$\underset{80\ gm}{\overset{50\ gm}{2NaOH}} + H_2SO_4 \longrightarrow \underset{142\ gm}{\overset{x\ gm}{Na_2SO_4}} + 2H_2O$$

$$\frac{50\ gm}{80\ gm} = \frac{x\ gm}{142\ gm}$$

$$x = \frac{50 \times 142}{80} = 88.75$$

∴ 88.75 gm of sodium sulfate would be formed.

13.

$$HCl + \underset{40\ gm}{\overset{75\ gm}{NaOH}} \longrightarrow \underset{58.5\ gm}{\overset{x\ gm}{NaCl}} + H_2O$$

$$\frac{75\ gm}{40\ gm} = \frac{x\ gm}{58.5\ gm}$$

$$x = \frac{\overset{15}{\cancel{75}} \times 58.5}{\underset{8}{\cancel{40}}} = 109.69$$

∴ 109.7 gm of sodium chloride would be formed.

14.

$$\underset{36.5\ lb}{\overset{x\ lb}{HCl}} + NaOH \longrightarrow \underset{58.5\ lb}{\overset{100\ lb}{NaCl}} + H_2O$$

$$\frac{x\ lb}{36.5\ lb} = \frac{100\ lb}{58.5\ lb}$$

$$x = \frac{100 \times 36.5}{58.5} = 62.4$$

62.4 lb of 100% hydrochloric acid are required.

Thus 624 lb of 10% hydrochloric acid are required.

15. 2.45% of 5kg may be expressed as .0245 x 5000 gm = 122.5 gm. Thus 122.5 gm of sodium chloride would be present.

16.
$$\underset{170\text{ gm}}{\overset{x\text{ gm}}{AgNO_3}} + \underset{58.5\text{ gm}}{\overset{122.5\text{ gm}}{NaCl}} \longrightarrow AgCl + NaNO_3$$

$$\frac{x\ \cancel{gm}}{170\ \cancel{gm}} = \frac{122.5\ \cancel{gm}}{58.5\ \cancel{gm}}$$

$$x = \frac{122.5 \times 170}{58.5} = 364.5$$

∴ 364.5 gm of silver nitrate would be required.

17.
$$AgNO_3 + \underset{58.5\text{ gm}}{\overset{122.5\text{ gm}}{NaCl}} \longrightarrow \underset{143.5\text{ gm}}{\overset{x\text{ gm}}{AgCl}} + NaNO_3$$

$$\frac{122.5\text{ gm}}{58.5\text{ gm}} = \frac{x\text{ gm}}{143.5\text{ gm}}$$

$$x = \frac{143.5 \times 122.5}{58.5} = 300.5$$

∴ 300.5 gm of silver chloride are produced.

UNIT 5: MECHANICS OF FLUIDS

CHAPTER 11. MOLECULAR FORCES IN LIQUIDS

Page 123

1. Molecules in gas being relatively farther apart have more freedom of motion.

2. Air surrounding the damp clothes does not become saturated with moisture since it is in motion.

3. Vaporization is often referred to a change from liquid to a gaseous state which takes place at the upper surface of a liquid. Boiling takes place throughout the liquid, usually at a definite temperature.

4. (a) The temperature of the liquid and the atmosphere.
   (b) The nature of the liquid.
   (c) The pressure upon the exposed surface of the liquid.
   (d) The degree of saturation of the air above the liquid.
   (e) The rate of circulation of the air over the surface of the liquid.
   (f) Surface area of the liquid.

5. A liquid which changes rapidly to a gas at ordinary temperatures.

6. (a) Attraction of like molecules.
   (b) Attraction of unlike molecules.

7. (a) The adhesive forces of water for glass are greater than the cohesive forces of water.

(b) The cohesive forces of mercury are greater than the adhesive forces of mercury for glass.

8. Acts as if it were elastic.

9. The needle is supported by the film (surface tension) which is formed by the cohesive forces of water.

Greasing the needle reduces the adhesive forces of water for steel.

10. Adhesion of water for glass is greater than cohesion of water molecules for each other.

The meniscus of a column of mercury is concave down because its cohesive forces are greater than the adhesive forces of mercury for glass.

11. Cohesive forces tend to make the drop occupy the least space.

12. The rise or fall of the liquid level in capillary tubes above or below the level outside the tubes.

- upon whether or not the liquid wets the tube.

Heat energy increases the kinetic energy of the molecules and so reduces the adhesive forces.

13. This depends upon the type of cultivation. If the cultivation increases the sizes of the spaces between the soil particles, this will prevent the moisture from rising to the surface by capillarity.

14. The sap has a greater density than the water in the soil. Osmotic pressure works from a lower density to a greater density, forcing the sap to rise.

Tiny openings are present in the organic fibers. Since the adhesive forces of sap for the fibers are greater than the cohesive forces, sap will rise through these openings by capillarity.

Evaporation causes a lower pressure area.

15. Yes. A greased strainer does not allow the strainer to get wet. The cohesive forces of water form surface tension which supports the mass of water.

16. Surface tension supports the insects.

17. Osmosis is the exchange which takes place between two fluids of different densities when separated by a moist semi-permeable membrane.

Osmostic pressure is caused by the fact that a lower density liquid displaces a higher density liquid at a greater rate.

18. Swelling of dried fruit; rising of sap and related plant processes.

19. The hot needle increases the kinetic energy of the water molecules. As the velocity of the molecules is increased, they move farther apart, pushing the tooth picks apart.

20. In a narrow tube each unit volume of the liquid comes in contact with a greater surface area of the container, causing the adhesive forces to increase more in relation to the gravitational force.

21. Rise of oil in a lamp wick.

Absorption of ink by a blotter.

Rise of sap in a tree.

Rise of water in soil.

CHAPTER 12. PRESSURE IN LIQUIDS AT REST

Page 133.

1. Force is the total push or pull exerted on a body.

   Pressure is a force per unit area.

2. "Head" is the vertical distance between the upper and the lower surfaces of a liquid.

3. (a) A hydraulic lift.
   (b) Hydraulic brakes.
   (c) Hydraulic press.
   (d) Pressure in water pipes.
   (e) Hydraulic air compressor.
   (f) Hydraulic elevator.
   (g) Hydraulic ram.

4. (a) The depth of the liquid.
   (b) The density of the liquid.

5. (a) Lb. wt. per square inch.
   (b) Lb. wt.

6. Liquid pressure is directly proportional to the depth, but independent of the area or shape of the container.

7. More pressure is exerted at the base of the dam because liquid pressure is directly proportional to the depth.

8. Pressure applied anywhere on a confined fluid is transmitted without loss throughout the body of the fluid.

9. (a) Shaping metals.
   (b) Hydraulic lifts.
   (c) Baling cotton.
   (d) Punching holes through steel plates.
   (e) Forging.

(f) Die-casting.

(g) Hydraulic brakes.

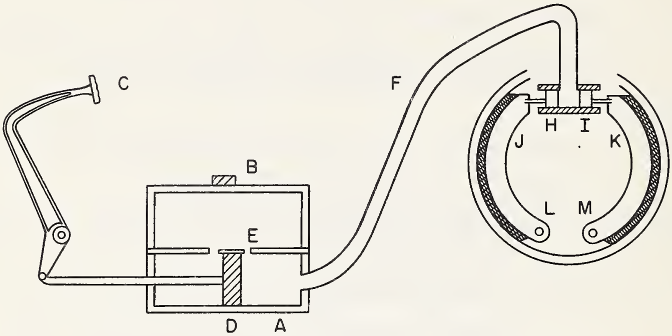
(h) Book-binding.

(i) Baling hay.

(j) Hydraulic valve lifters.

10. Velocity, or distance through which the force acts.

11.



The master cylinder consists of the cylinder proper A and the supply tank B immediately above it. The supply tank maintains a constant volume of the brake fluid. When the pedal C is pressed the piston D moves to the right, closing the opening E and forcing fluid from the cylinder through the tubing F into the wheel cylinder which contains two pistons H and I connected by links to the brake shoes J and K which are pivoted at L and M.

When the liquid is forced into the wheel cylinder the pistons move outward, pushing the shoes J and K against the inner surface of the brake drum.

12. Force ratio is the weight supported (or the resistance overcome) divided by the force applied.

Page 134 Problems.

1. $p = hd$

$= 200 \text{ ft} \times 62.5 \frac{\text{lb. wt.}}{\text{ft}^3}$

$= 12,500 \frac{\text{lb. wt.}}{\text{ft}^2}$

2. $F = Ahd$

$= 2 \text{ ft}^2 \times 1 \text{ ft} \times 62.5 \frac{\text{lb. wt.}}{\text{ft}^3}$

$= 125$ lb. wt.

3. $p = hd$

$= 10 \text{ ft} \times 62.5 \frac{\text{lb. wt.}}{\text{ft}^3}$

$= 625 \frac{\text{lb. wt.}}{\text{ft}^2}$

4. $F = Ahd$

$= 6 \text{ m}^2 \times 1 \text{ m} \times 1000 \frac{\text{kg. wt.}}{\text{m}^3}$

$= 6000$ kg. wt.

5. $p = hd$

$15 \frac{\text{lb. wt.}}{\text{in.}^2} = h \times 62.5 \frac{\text{lb. wt.}}{1728 \text{ in.}^3}$

$h = 15 \frac{\text{lb. wt.}}{\text{in.}^2} \times \frac{1728 \text{ in.}^3}{62.5 \text{ lb. wt.}} = 414.7$ in. or 34.6 ft.

6. $p = hd$

$= 5 \text{ ft} \times 62.5 \frac{\text{lb. wt.}}{\text{ft}^3}$

$= 312.5 \frac{\text{lb. wt.}}{\text{ft}^2}$

7. $p = hd$

$= 72 \text{ ft} \times 62.5 \frac{\text{lb. wt.}}{\text{ft}^3}$

$= 4500 \frac{\text{lb. wt.}}{\text{ft}^2}$

8. $p = hd$

$30 \frac{\text{lb wt.}}{\text{in}^2}$ or $30 \times 144 \frac{\text{lb. wt.}}{\text{ft}^2} = h \times 62.5 \frac{\text{lb. wt.}}{\text{ft}^3}$

$h = 30 \times 144 \frac{\text{lb. wt.}}{\text{ft}^2} \times \frac{1}{62.5} \frac{\text{ft}^3}{\text{lb. wt.}} = 69.1$ ft.

9. (a) $F = Ahd$

$= 800 \text{ ft}^2 \times 8 \text{ ft} \times 62.5 \frac{\text{lb. wt.}}{\text{ft}^3}$

$= 400{,}000$ lb. wt.

(b) Area of one end $= 20 \text{ ft} \times 8 \text{ ft} = 160 \text{ ft}^2$

$F = \frac{1}{2} Ahd$

$= \frac{1}{2} \times 160 \text{ ft}^2 \times 8 \text{ ft} \times 62.5 \frac{\text{lb. wt.}}{\text{ft}^3} = 40{,}000$

10. $F = Ahd$

$= 2 \text{ ft}^2 \times 500 \text{ ft} \times 62.5 \frac{\text{lb. wt.}}{\text{ft}^3}$

10. (Cont'd)

$F = 62{,}500$ lb. wt.

11. (a) $p = hd$

$= 80 \text{ cm} \times 1 \frac{\text{gm. wt.}}{\text{cm}^3} = 80 \frac{\text{gm. wt.}}{\text{cm}^2}$

(b) $p = hd$

$= 80 \text{ cm} \times 0.80 \frac{\text{gm. wt.}}{\text{cm}^3} = 64 \frac{\text{gm. wt.}}{\text{cm}^2}$

12. $p = hd$

$= 290 \text{ ft} \times 64 \frac{\text{lb. wt.}}{\text{ft}^3} = 18{,}560 \frac{\text{lb. wt.}}{\text{ft}^2}$

13. Pressure on large piston = pressure on small piston.

$$\frac{X}{\pi (20)^2 \text{in}^2} = \frac{500 \text{ lb. wt.}}{\pi (1)^2 \text{in}^2}$$

$$X = \frac{500 \text{ lb. wt.}}{\pi \text{ in}^2} \times \pi \times 400 \text{ in}^2 = 200{,}000 \text{ lb.wt.}$$

or

Since the area of the large piston is 400 times that of the small piston the force ratio is 400. Therefore the force is multiplied by 400, i.e. 500 lb. wt. x 400 = 200,000 lb. wt.

14. Area of piston $= \pi r^2 = 3.142 \times 9 \text{ in}^2$.
Force on this area = 3142 lb. wt.
Pressure $= \frac{F}{A} = \frac{3142 \text{ lb.wt.}}{31.42 \times 9 \text{ in}^2} = \frac{1000 \text{ lb.wt.}}{9 \text{ in.}^2} = 111.1 \text{ lb.wt./in}^2$

15. Since the diameter is 20 times as great, the area is 400 times as great. The force ratio is 400,

or $\text{F.R.} = \frac{R}{F} = \frac{D^2}{d^2} = \frac{20^2}{1^2} = 400.$

16. Force on small piston $= \frac{3 \text{ tons wt.}}{100}$

or $\frac{6000 \text{ lb. wt.}}{100} = 60$ lb. wt.

17. $\text{F.R.} = \frac{D^2}{d^2} = \frac{(10 \text{ in})^2}{(1 \text{ in})^2} = 1000$

18. (a) $\text{F.R.} = \frac{D^2}{d^2} = \frac{(30 \text{ in})^2}{(2 \text{ in})^2} = 225.$

(b) Effort exerted $= 15 \times \frac{2000 \text{ lb. wt.}}{225} = 133$ lb. wt.

19. (a) $\text{F.R.} = \frac{D^2}{d^2} = \frac{30 \text{ in}^2}{2 \text{ in}^2} = 15$

(b) Weight raised = 15 lb. wt. x 15 = 225 lb. wt.

$2 \text{ in}^2 \times 10 \text{ in} = 30 \text{ in}^2 \times X$

$\therefore X = 0.67$ in.

20. $F.R. = \frac{R}{F} = 4 \times \frac{2000 \text{ lb. wt.}}{20 \text{ lb. wt.}} = 400$

$$F.R. = \frac{D^2}{d^2}$$

$$400 = \frac{(20 \text{ in})^2}{d^2}$$

$$d^2 = \frac{(20 \text{ in})^2}{400} = \frac{400 \text{ in}^2}{400} = 1 \text{ in}^2$$

$$d = \sqrt{1 \text{ in}^2} = 1 \text{ in}$$

CHAPTER 13. PRESSURE IN GASES

Page 148.

Questions.

1. Troposphere extends to a height of about 7 miles above the earth's surface. Region in which clouds are formed.

Stratosphere - region above the troposphere. It extends to about 50 miles above the earth's surface. No cloud formation is present in this region.

Tropopause separates the stratosphere from the troposphere.

Chemosphere is the upper part of the stratosphere. Region in which most of the meteors burn out and auroras are formed.

Ionosphere stretches to about 250 miles. Temperature increases rapidly in this region. Reflects radio waves.

Exosphere - region above the ionosphere. Very little air present.

2. Troposphere.

3. Reflects radio waves.

4. Weight of 1 liter of air is 1.29 gm. wt.

5. Pressure of the atmosphere is equal to the pressure of the fluids within our bodies.

6. Find the cross-sectional area in square inches and multiply by 14.7 lb. wt. per sq. in.

7. Atmospheric pressure on the surface of the liquid forces it up until a point is reached where the gravitational force on the water is equal to the atmospheric pressure.

8. Atmospheric pressure holds the water in the jug.

9. The same. The height of mercury does not depend upon the shape or the size of the tube. The greater the diameter of the tube, the greater the surface area upon which the atmospheric pressure acts.

10. Upon expansion of the lungs a low pressure area is formed within the lungs. The higher pressure of the atmosphere forces air into the lungs.

11. A tire gauge indicates the excess of air pressure within the tire above that of the atmosphere. Since the atmospheric pressure on top of the mountain is less than that below, the reading of the gauge is higher on top of the mountain.

12. (a) 14.7 lb. wt./sq. in.
(b) 1034 gm. wt./$cm.^2$
(c) 76 cm. of mercury

13. Advantages:

Great accuracy.

13. (Cont'd) Disadvantages:

(a) Must be corrected for changes in temperature.
(b) Inconvenient to handle.
(c) Mercury responds slowly to pressure changes.

14. Advantages:

(a) Very sensitive to changes in pressure.
(b) Portable.

Disadvantage:

Not as accurate as the mercury barometer.

15. A barograph is a self-recording barometer. It is used to record barometric readings on paper.

16. The temperature and pressure of the gas must be specified.

17. The volume of a given mass of gas varies inversely as the pressure applied, if the temperature remains constant.

18. The greater the pressure on a given mass of gas, the greater its density because its volume is reduced.

19. Volumes of given masses of gases vary with changes in pressure, but liquids are incompressible.

20. $V_1P_1 = V_2P_2$ or $VP = K$.

Page 148. Problems.

1. Measure length, width and height of the room.
Use $V = lwh$.. (ft^3).

2. Volume of room $= \dfrac{V\ ft^3}{27\ \dfrac{ft^3}{yd^3}} = \dfrac{V}{27}\ yd^3$

Mass of air in room = V x D.

$= \dfrac{V}{27}\ yd^3 \times \dfrac{2\ lb}{yd^3} = \dfrac{2}{27} V\ lb.$

3. Mass of air in flask = 2.5 gm.
Volume of air is 2000 cc. or 2 liters.
Mass of 1 liter of air = 1.25 gm.

4. Area of cross-sectional surface of hemisphere
$= 2\pi\ (11\ in.)^2$
$= 761\ in.^2$

Force required $= 761\ in.^2 \times \dfrac{15\ lb.\ wt.}{in.^2}$

$= 11{,}415$ lb. wt.

5. Mercury reading is 74 cm.

(a) Since water is $\dfrac{1}{13.7}$ times as heavy as mercury, the reading of a water barometer is 13.7 times that of the mercury reading.

Water reading = 74 cm. x 13.7 = 1014 cm.

(b) Density of glycerine = $\frac{1.25 \text{ gm. wt.}}{\text{c.c.}}$

That is, glycerine is 1.25 times as heavy as water.
Glycerine reading = $\frac{1014 \text{ cm.}}{1.25}$ = 811 cm.

6. Cross-sectional area of hemisphere = $2\pi(1 \text{ in.})^2$ = 6.3 in.^2

Barometric pressure = 29 in. of mercury.

Density of mercury = $\frac{849 \text{ lb. wt.}}{\text{ft}^3}$ (p. 139)

or $\frac{849 \text{ lb. wt.}}{1728 \text{ in.}^3}$

.'. the pressure = 29 in. x $\frac{849}{1728}\frac{\text{lb.wt.}}{\text{in.}^3}$ = 14.2 $\frac{\text{lb.wt.}}{\text{in.}^2}$

or pressure = $\frac{29}{30}$ x 14.7 $\frac{\text{lb.wt.}}{\text{in.}^2}$ = 14.2 $\frac{\text{lb.wt.}}{\text{in.}^2}$

.'. Force required = 14.2 $\frac{\text{lb.wt.}}{\text{in.}^2}$ x 6.3 in.^2 = 89.5 lb.wt.

Pressure inside is 9.7 in. of mercury.

Difference in pressure = 29 in. - 9.7 in. = 19.3 in. of mercury.

This difference is equivalent to $\frac{19.3}{30}$ x 14.7 $\frac{\text{lb.wt.}}{\text{in.}^2}$

= 9.5 $\frac{\text{lb. wt.}}{\text{in.}^2}$

.'. Force required = 9.5 $\frac{\text{lb.wt.}}{\text{in.}^2}$ x 6.3 in.^2 = 60 lb.wt.

7. 1 atmosphere = 14.7 $\frac{\text{lb.wt.}}{\text{in.}^2}$

3 atmospheres = 14.7 $\frac{\text{lb.wt.}}{\text{in.}^2}$ x 3 = 44.1 $\frac{\text{lb.wt.}}{\text{in.}^2}$

8. Pressure = 735 mm. or 73.5 cm.

73.5 cm. x 13.6 $\frac{\text{gm}}{\text{cm}^3}$ = 1000 $\frac{\text{gm.wt.}}{\text{cm}^2}$

9. Total surface area = 2(24 in.^2) + 2(12 in.^2) + 2(18 in.^2)

= 108 in.^2

Pressure = 108 in.^2 x $\frac{15 \text{ lb.wt.}}{\text{in.}^2}$ = 1620 lb.wt.

10. Pressure of water (only) = 15 ft. x 62.5 $\frac{\text{lb.wt.}}{\text{ft.}^3}$ =

937.5 $\frac{\text{lb.wt.}}{\text{ft.}^2}$

(continued)

10. (Cont'd) or $\frac{6.5 \text{ lb.wt.}}{\text{in.}^2}$

Pressure of air (only) = $14.7 \frac{\text{lb.wt.}}{\text{in.}^2}$

Total pressure (water and air) = $6.5 \frac{\text{lb.wt.}}{\text{in.}^2} + 14.7 \frac{\text{lb.wt.}}{\text{in.}^2}$

$= 21.2 \frac{\text{lb.wt.}}{\text{in.}^2}$

11. Pressure of water (only) = $3000 \text{ cm.} \times 1 \frac{\text{gm.wt.}}{\text{cm.}^3} = 3000 \frac{\text{gm.wt.}}{\text{cm.}^2}$

Pressure of air (only) = $76 \text{ cm.} \times 13.6 \frac{\text{gm.wt.}}{\text{cm.}^3} = 1034 \frac{\text{gm.wt.}}{\text{cm.}^2}$

Combined pressure = 4034 gm. wt./cm.^{2}

12. Pressure of water (only) = $150 \text{ ft.} \times 62.5 \frac{\text{lb.wt.}}{\text{ft.}^3} = 9375 \frac{\text{lb.wt.}}{\text{ft.}^2}$

or $65 \frac{\text{lb.wt.}}{\text{in.}^2}$

Pressure of air = $14.7 \frac{\text{lb. wt.}}{\text{in.}^2}$

Combined pressure = $79.7 \frac{\text{lb.wt.}}{\text{in.}^2}$ or $80 \frac{\text{lb.wt.}}{\text{in.}^2}$

Page 149.

13. Mercury reading = 760 mm.

(a) Water reading = 760 mm. x 13.6 = 10336 mm. or 1034 cm.

(b) Alcohol reading = $\frac{1034}{0.80}$ cm. = 1293 cm.

14. $V_1P_1 = V_2P_2$

$V \times 1 \text{ at.} = 10 \text{ ft.}^3 \times 2 \text{ at.}$

$V = \frac{10 \text{ ft.}^3 \times 2 \text{ at.}}{1 \text{ at.}} = 20 \text{ ft.}^3$

15. $V_1P_1 = V_2P_2$

$V \times 1 \text{ at.} = 10 \text{ l.} \times 2.5 \text{ at.}$

$V \times \frac{10 \text{ l.} \times 2.5 \text{ at.}}{1 \text{ at}} = 25 \text{ l.}$

For the above solution it is assumed that the 2.50 at. pressure includes atmospheric pressure.

16. $V_1P_1 = V_2P_2$

$400 \text{ cm.}^3 \; P = 1000 \text{ cm.}^3 \times 76.0 \text{ cm.}$

$P = \frac{1000 \text{ cm.}^3 \times 76.0 \text{ cm.}}{400 \text{ cm.}^3} = 190 \text{ cm.}$

17. $V_1P_1 = V_2P_2$

$100 \text{ ft.}^3 \times P = 20 \text{ ft.}^3 \times 5 \text{ at.}$

$$P = \frac{20 \text{ ft.}^3 \times 5 \text{ at.}}{100 \text{ ft.}^3} = 1 \text{ at.}$$

18. $V_1P_1 = V_2P_2$

$V \times 800 \text{ mm.} = 500 \text{ ml.} \times 750 \text{ mm.}$

$$V = \frac{500 \text{ ml.} \times 750 \text{ mm.}}{800 \text{ mm.}} = 469 \text{ ml.}$$

19. Difference in pressure = 0.4 in. of mercury
or 0.4 in. x 13.6 = 5.44 in. of water

Density of air = $1.29 \dfrac{\text{gm.}}{\text{liter}}$

Density of water = $1000 \dfrac{\text{gm.}}{\text{liter}}$

Specific Gravity of air = 0.00129.

$\therefore$ Difference in pressure = $\dfrac{5.44 \text{ in.}}{0.00129}$ of air

or $\dfrac{5.44}{12 \times 0.00129}$ ft. of air

= 351 ft. of air.

$\therefore$ Height of hill = 336 ft.

20. Density of sulphur dioxide = $\dfrac{64 \text{ gm.}}{1.}$ at 760 mm. and 0° C.

64 gm. at 760 mm. and 0°C. occupy 1 l.

Use Boyle's Law to find volume.

64 gm. would occupy at 800 mm. and 0°C.

$V_1P_1 = V_2P_2$

$V \times 800 \text{ mm.} = 1 \text{ l.} \times 760 \text{ mm.}$

$$V = \frac{1 \text{ l.} \times 760 \text{ mm.}}{800 \text{ mm.}} = 0.95 \text{ l.}$$

$\therefore$ 0.95 l. at 800 mm. has a mass of 64 gm.

Density at 800 mm. and 0°C = $\dfrac{64 \text{ gm.}}{0.95 \text{ l.}} = 67.4 \dfrac{\text{gm.}}{\text{l.}}$

21. Difference in pressure = 10 cm. of mercury.

Density of mercury = $13.6 \dfrac{\text{gm.wt.}}{\text{cm}^3}$

Density of air = $0.00123 \dfrac{\text{gm.wt.}}{\text{cm}^3}$

(continued)

21. (Cont'd)

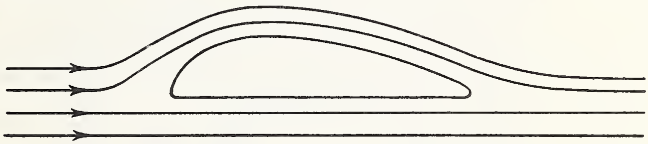
That is, mercury = $\dfrac{13.6\ \frac{\text{gm.wt.}}{\text{cm}^3}}{0.00123\ \frac{\text{gm.wt.}}{\text{cm}^3}}$

or 11,057 times as heavy as air.

$\therefore$ Height of airplane = 10 cm. x 11,057 = 110,570 cm. or 1106 m.

CHAPTER 14. FLUIDS IN MOTION

Page 153. Questions.

1. Pressure is directly proportional to depth.. Pressure is directly proportional to density.
2. Water flowing in the water pipes reduces the pressure.
3. The greater the flow the smaller the pressure or pressure varies inversely as the rate of flow.
4. 

   Because of the shape of the wing, air moves at a greater velocity on top than at the bottom. This creates a lower pressure at the top and a greater pressure at the bottom.
5. Because of the spin of the ball, the velocity of air past its surface is greater on one side than the other. The difference in air pressure causes it to curve.
6. To the right. Velocity of air with respect to the ball is greater on the left.
7. The velocity of the jet is directly proportional to the depth of the liquid.
8. The wind blowing across the top of the chimney creates a low pressure area (Bernoulli's Principle). The higher pressure at the bottom of the chimney forces the air up.
9. The efficiency of a vacuum cleaner depends upon the difference in pressure between the inside of the cleaner and the outside.
   This difference is created by a vacuum pump.
10. As the air moves past the upper surface of the paper, a low pressure area is created (Bernoulli's Principle). The higher pressure at the bottom of the paper causes it to rise.
11. During the busy time of the day when a lot of water is used, the flow of water through the pipes decreases the pressure. A water reservoir increases the pressure by the force of gravity.

UNIT 6: HEAT

CHAPTER 16. HEAT AND TEMPERATURE

Page 167. Questions.

1. The difference in the intensity of heat possessed by two bodies causing heat to flow from the hotter body to the colder body.

2. Hot - the presence of heat energy.
Cold - the absence of heat energy.

3. Temperature refers to the intensity of heat.
Quantity of heat is the amount of heat energy possessed by a body.

4. (a) Awkward to handle.

(b) Cannot be used for readings above $100^{\circ}C$. Water boils at $100^{\circ}C$ under standard pressure.

(c) Cannot be used for readings below $0^{\circ}C$. Water freezes at $0^{\circ}C$ under standard pressure.

(d) Water evaporated quite readily increasing the pressure in the bulb.

(e) Water level changes with changes in atmospheric pressure.

5. Mercury is quite sensitive to temperature changes. It does not evaporate readily. However, it has a fairly high freezing point and therefore cannot be used to measure low temperatures. Mercury has a fairly high boiling point and can be used to measure high temperatures.

6. Alcohol has a low freezing point and can be used for measuring low temperatures. However it boils quite readily and therefore cannot be used for measuring high temperatures.

7. When heated, a body

(a) shows an increase in temperature

(b) expands (a few exceptions)

(c) changes its state.

Page 167. Problems.

1. (a) $\dfrac{t - 32F^{\circ}}{180F^{\circ}} = \dfrac{0^{\circ}C - 0C^{\circ}}{100C^{\circ}}$

$t - 32F^{\circ} = 0$

$t = 32F^{\circ}$ above zero

or $t = 32^{\circ}F$

(b) $20^{\circ}C = 68^{\circ}F$

(c) $100^{\circ}C = 212^{\circ}F$

(d) $\frac{t - 32F^o}{180F^o} = \frac{50^oC - 0C^o}{100C^o}$

$t - 32F^o = \frac{50C^o}{100C^o} \times 180F^o$

$t - 32F^o = 90F^o$

$t = 122F^o$ above zero

or 122^oF

or

$F = \frac{9}{5} C + 32$

$F = \frac{9}{5} \times 50 + 32 = 122$

$\therefore \; 50^oC = 122^oF.$

(e) $75^oC = 167^oF$

(f) $-40^oC = 40^oF$

(g) $\frac{t-32F^o}{180F^o} = \frac{-195^oC - 0C^o}{100C^o}$

$\frac{t-32F^o}{180F^o} = \frac{-195C^o}{100C^o}$

$t - 32F^o = \frac{-195C^o}{100C^o} \times 180F^o$

$t - 32F^o = -351F^o$

$t = -319F^o$ above zero

$t = -319^oF$

or

$F = \frac{9}{5} C + 32$

$= \frac{9}{5} (-195) + 32$

$= -351 + 32 = -319$

$\therefore \; t = -319^oF$

2. (a) $\frac{32^oF - 32^oF}{180F^o} = \frac{t - 0C^o}{100C^o}$

$t - 0C^o = 0$

$t = 0C^o$ or 0^oC.

or $C = \frac{5}{9}(F - 32)$

$C = \frac{5}{9}(32 - 32)$

$C = 0$

$\therefore\ t = 0^{o}C.$

(b) $68^{o}F = 20^{o}C$

(c) $212^{o}F = 100^{o}C$

(d) $50^{o}F = 10^{o}C$

(e) $\frac{t - 0C^{o}}{100C^{o}} = \frac{95^{o}F - 32F^{o}}{180F^{o}}$

$t - 0^{o}C = \frac{63F^{o}}{180F^{o}} \times 100C^{o} = 35C^{o}$

$t = 35^{o}C$ above 0 or $35^{o}C$

or $C = \frac{5}{9}(F - 32)$

$C = \frac{5}{9}(95 - 32)$

$C = \frac{5}{9}(63) = 35$

$\therefore\ t = 35^{o}C$

(f) $-40^{o}F = -40^{o}C.$

3. Let Fahrenheit reading be X

Then Centigrade reading is X

Using $F = \frac{9}{5}C + 32$

$X = \frac{9}{5}X + 32$

$-\frac{4}{5}X = 32$

$X = 32\left(\frac{-5}{4}\right) = -40$

$\therefore$ two readings are the same at -40^{o}.

4. Temperature change $= 60^{o}C - 20^{o}C = 40C^{o}$

But $40C^{o} = 40C^{o} \times \frac{9}{5}\frac{F^{o}}{C^{o}} = 72F^{o}$

5. $1C^{o} = \frac{9}{5}F^{o}$

$5C^{o} = \frac{9}{5} \times 5F^{o} = 9F^{o}$

$15C^{o} = \frac{9}{5} \times 15F^{o} = 27F^{o}$

$45C^{o} = \frac{9}{5} \times 45F^{o} = 81F^{o}$

$65C^{o} = \frac{9}{5} \times 65F^{o} = 117F^{o}$

6. $1\ F^{o} = \frac{5}{9}\ C^{o}$

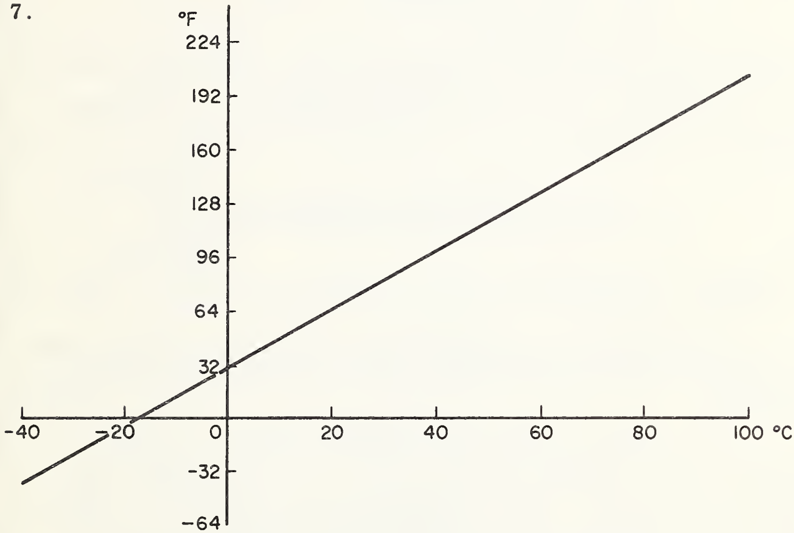
$9F^{o} = \frac{5}{9} \times 9C^{o} = 5C^{o}$

$27F^{o} = \frac{5}{9} \times 27C^{o} = 15C^{o}$

$36F^{o} = \frac{5}{9} \times 36C^{o} = 20C^{o}$

$95F^{o} = \frac{5}{9} \times 95C^{o} = 52.8C^{o}$

7.



CHAPTER 17. EXPANSION OF SOLIDS

Page 117. Questions.

1. Heat increases the kinetic energy of the atoms, causing them to vibrate with greater average speed and push one another away.

2. Increase per unit length per unit degree rise in temperature.

3. (a) Increase in unit area per degree rise in temperature.

(b) Increase in unit volume per degree rise in temperature.

4. The metallic ring has a higher coefficient of expansion than the glass jar.

5. A tumbler made of thin glass is less likely to crack when hot water is poured into it because the heat is distributed more uniformly causing a uniform expansion. In thick glass one surface expands more rapidly than the other side. This uneven pressure causes it to crack.

6. The hole becomes larger. Consider the plate with the hole as a ring; its diameter increases. Or, consider the plate without the hole; the plate expands. Now consider the circular part cut out; it also expands upon heating.

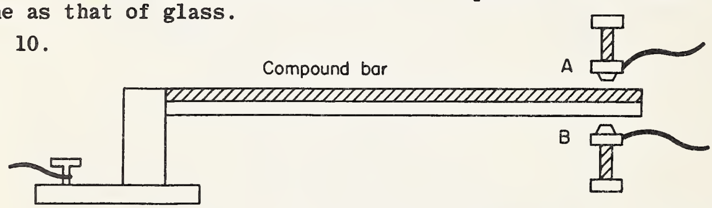
7. (a) Two metals which have different rates of expansion riveted together. When heated, one expands at a greater rate than the other causing bending.

(b) (1) Oven thermostats.
(2) Furnace thermostats.
(3) Fire alarm.

8. The length of the pendulum expands. The period of the pendulum varies directly as the square root of the length.

9. Platinum. Its coefficient of expansion is almost the same as that of glass.

10.



The bending of a compound bar when heated is applied in the construction of thermostats, which are used mainly for controlling the temperature of buildings heated by hot-air furnaces or boilers.

The essential parts of a thermostat are a compound bar and two electric contact points A and B. When the temperature rises to a certain degree, the compound bar bends down to make contact with point B. This completes an electric circuit which shuts off the furnace. When the temperature drops to a certain degree, the compound bar straightens out or bends up to make contact with point B. This completes an electric circuit which opens the valve of the furnace.

Page 172. Problems.

1. $1\ C^{o} = \frac{5}{9}\ F^{o}$

$$0.000019/C^{o} = 0.000019 \times \frac{5}{9}/F^{o}$$

$$= 0.000011/F^{o}$$

2. (a) $0.000017/C^o \times 2 = 0.000034/C^o$
(b) $0.000017/C^o \times 3 = 0.000051/C^o$

3. $\Delta 1 = 1(t_2 - t_1)\alpha$

$= 10 \text{ ft.} \times 5C^o \times 0.000013/C^o$

$= 0.00065 \text{ ft.}$

4. $\Delta 1 = 1(t_2 - 5_1)\alpha$

$= 100 \text{ cm} \times 20C^o \times 0.000019/C^o$

$= 0.038 \text{ cm.}$

5. $\alpha = \dfrac{\Delta 1}{1\Delta t}$

$= \dfrac{0.189 \text{ cm.}}{100 \text{ cm.} \times 100C^o}$

$= 0.0000189/C^o$

6. $t_2 - t_1 = \dfrac{\Delta t}{1 + \alpha}$

$t_2 - 0^oC = \dfrac{0.022 \text{ cm.}}{100 \text{ cm} \times 0.000013/C^o}$

$t_2 - 0^oC = 19C^o$

i.e. t_2 is $19C^o$ above 0^oC

$\therefore t_2 = 19C^o$

7. Coef. of volume expansion of steel $= 0.000039/C^o$
$\Delta V = V \times \Delta t \times$ coef. of volume exp.

$= 1000 \text{ cm}^3 \times 100C^o \times 0.000039/C^o$

$= 3.9 \text{ cm.}^3$

8. $1 = 1 \times \Delta t \times \alpha$

$= 14 \text{ ft.} \times 180C^o \times 0.000013/C^o$

$= 0.033 \text{ ft.}$

9. $\Delta 1 = 1 \times \Delta t \times \alpha$

$= 10\text{m.} \times 50C^o \times 0.000013/C^o$

$= 0.0065 \text{ m.}$

10. $\alpha = \dfrac{\Delta 1}{1 \times \Delta t}$

$= \dfrac{0.20 \text{ in.}}{40 \times 12 \text{ in.} \times 90F^o} = 0.0000046/F^o$

or $0.0000083/C^o$

11. $\Delta 1 = 1 \times \Delta t \times \alpha$

$\Delta 1 = 1 \text{ in.} \times 80C^o \times 0.000013/C^o$

$= 0.000104 \text{ in.}$

12. Expansion of steel tape = l x Δt x α

= 100 ft. x $20C^{o}$ x $0.000013/C^{o}$

= 0.026 ft.

Error = 0.026 ft. in 100 ft.

or 0.013 ft. in 50 ft.

The distance measured off was actually 0.013 ft. too long.

CHAPTER 18. EXPANSION OF LIQUIDS

Page 175. Questions.

1. The force of attraction between the molecules in a liquid is less than that in a solid.

2. Water has a maximum density at $4^{o}C$. That is, as water is heated above $4^{o}C$. it expands and as it is cooled from $4^{o}C$ to $0^{o}C$ it expands.

Water contracts as it is heated from $0^{o}C$ to $4^{o}C$.

3. Water has a maximum density at $4^{o}C$. Water at $0^{o}C$ is lighter than water at $4^{o}C$ and is forced to the surface where it freezes at $0^{o}C$.

4. The dilatometer, being exposed to the heat, expands before the liquid inside it is heated. Thus the volume of the dilatometer is increased through expansion before the volume of the liquid inside it is increased.

5. Coefficient of real expansion is the rate at which the liquid expands regardless of its container.

Coefficient of apparent expansion is the rate at which the liquid appears to expand within a container because of the expansion of the container.

Coefficient of real expansion is equal to the coefficient of apparent expansion plus the coefficient of expansion of the container.

$$C_r = C_a + C_c$$

Page 175. Problems.

1. $1\ F^{o} = \frac{5}{9}\ C^{o}$

$$0.0014/C^{o} = 0.0014/C^{o} \times \frac{5}{9}\ C^{o}/F^{o}$$

$$= 0.0008/F^{o}$$

2. Coef. of linear expansion of glass = $0.0000090/C^{o}$ (p. 169)

Coef. of volume expansion of glass = $0.0000090/C^{o}$ x 3

= $0.000027/C^{o}$

Coef. of real expansion of mercury = $0.00018/C^{o}$ (Table 8)

$$C_a = C_r - C_c$$

2. (Cont'd) $C_a = 0.00018/C^o - 0.000027/C^o$

$= 0.000153/C^o$

Overflow is the apparent expansion.

$\Delta V = V \times \Delta t \times C_a$

$= 1000\ cm^3 \times 30C^o \times (0.0001800.000027)/C^o$

$= 1000\ cm^3 \times 30C^o \times (.000153/C^o)$

$= 4.59\ cm^3$

3. The increase in volume of alcohol in the tank is the real expansion of alcohol.

$\Delta V = V \times \Delta t \times C_r$

$V = 10$ gal. $\times 10\ C^o \times 0.0011/C^o$

$= 0.11$ gal.

4. Decrease in volume is the real contraction of alcohol.

A change of ($90^oF - 40^oF$) or $50F^o$ is the same as

$50F^o \times \frac{5\ C^o}{9\ F^o}$ or $28\ C^o$

$\Delta V = V \times \Delta t \times C_r$

$= 20\ ft^3 \times 28C^o \times 0.0011/C^o$

$= 0.62\ ft.^3$

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5. Overflow is equal to the apparent expansion.

$C_a = C_r - C_c$

$= 0.00096/C^o - 0.000013 \times 3/C^o = 0.00092/C^o$

$\Delta V = V \times \Delta t \times C_a$

$= 16$ gal. $\times 10\ C^o \times 0.00092/C^o$

$= 0.15$ gal.

6. Overflow is the apparent expansion. Coefficient of apparent expansion of turpentine in copper is equal to the coefficient of real expansion of turpentine minus the coefficient of volume expansion of copper.

Coef. of linear expansion of copper = $0.0000017/C^o$ (p. 169).

Coef. of volume expansion of copper = $0.000017/C^o \times 3$

$= 0.000051/C^o$

$C_a = C_r - C_c$

$= 0.00097/C^o - 0.000051/C^o$

$= 0.00092/C^o$

6. $\Delta V = V \times \Delta t \times C_a = 100 \text{ gal.} \times 25C^o \times 0.00092/C^o$

$= 2.3$ gal.

7. Change in volume of alcohol is the real expansion of alcohol.

$\Delta V = V \times \ t \times C_r$

$= 10 \text{ gal.} \times 25\ C^o \times 0.0011/C^o = 0.28$ gal.

$\therefore$ volume is reduced by 0.28 gal.

8. (a) The increase in volume reading is the apparent expansion of mercury in pyrex glass.

The apparent expansion of mercury = 0.144 c.c.

(b) Coef. of apparent expansion =

$$\frac{\text{apparent expansion}}{\text{original volume} \times \text{temperature change}}.$$

$$C_a = \frac{\Delta V_a}{V \times \Delta t}$$

$$= \frac{0.144 \text{ c.c.}}{10.000 \text{ c.c.} \times 80C^o} = 0.00018/C^o.$$

(c) $C_r = C_a + C_c$

$= 0.00018/C^o + 0.000004 \times 3/C^o$

$= 0.000192/C^o$

(d) Absolute or real expansion of mercury = 10.000 c.c. $\times\ 0.000192/C^o = 0.154$ c.c.

9. $\Delta V = V \times \Delta t \times C_r$

$= 1 \text{ gal.} \times 30\ C^o \times 0.00124/C^o$

$= 0.037$ gal.

10. Increase in volume of mercury $= 1 \text{ c.c.} \times 80\ C^o \times 0.00018/C^o$

$= 0.0144$ c.c.

New volume = 1.0144 c.c.

Mass of 1.0144 c.c. of mercury at 100^oC = 13.55 gm.

Mass of 1.0 c.c. of mercury at $100^oC = \frac{13.55 \text{ gm.}}{1.0144} = 13.36$ gm.

11. Volume of mercury at 0^oC = 0.5 c.c.

Apparent increase in volume of mercury =

$0.5 \text{ c.c.} \times 100C^o \times 0.0015/C^o = 0.075$ c.c.

But the apparent expansion of mercury is equal to the volume of the stem between the 0^oC and the 100^oC marks.

Volume of stem $(\pi r^2 h) = 3.143\ (0.01\text{cm})^2\ h$

$= 0.003143\ h\ \text{cm}^2$

$\therefore$ $0.0003143\ h\ cm^2 = 0.075\ cm^3$

$h = 238$ cm.

$\therefore$ distance between $0^{o}C$ and $100^{o}C$ marks is 238 cm.

Note: The coefficient of apparent expansion of mercury in this thermometer is given as 0.0015/C°. It should read 0.00015/C°. This will give an answer of 23.8 cm. for the distance between the 0°C and the 100°C marks.

CHAPTER 19. EXPANSION OF GASES

Page 182. Questions.

1. A given mass of gas refers to a definite amount or a specific number of molecules. During the experiments no molecules are added to or subtracted from this number.

2. When the temperature of a gas is increased, its thermal energy is also increased and so the velocity of the molecules becomes greater causing them to move farther apart.

3. (a) "Ideal" refers to a gas which would react uniformly to all changes in temperature and pressure as the mass of the molecules and the intermolecular attraction are zero.

(b) Absolute zero refers to the temperature at which all energy has been completely removed. At -273°C the molecules possess absolutely no energy. The molecules are motionless.

4. The increase in volume experienced by a unit volume of a gas when its temperature is raised one centigrade degree. This increase is compared to the volume the gas would occupy at 0°C. The coefficient of expansion of any gas is equal to 1/273 of the volume it occupies at 0°C.

5. Heat is a form of energy. As the gas is heated, its molecules are energized and their velocity is increased, bombarding the walls of the container at a greater rate, resulting in an increase in pressure.

6. $V_1P_1 = V_2P_2$ or $VP = K$

(if $T = K$)

7. $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ or $\frac{V}{T} = K$ (if $P = K$ and T is expressed in absolute degrees)

8. The volume of a given mass of gas varies directly as the absolute temperature and inversely as the pressure.

$$\frac{V_1P_1}{T_1} = \frac{V_2P_2}{T_2} \quad \text{or} \quad \frac{VP}{T} = K.$$

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9. (a) If the absolute temperature of a given mass of gas is doubled, the volume is doubled.

9. (b) Its mass remains constant.

(c) Its density is decreased to one-half since the mass remains constant and the volume is doubled.

10. Volumes of gases change readily with changes in temperature and pressure, therefore if we wish to compare volumes of gases we must reduce the temperature and pressure to some standard. Remember that equal volumes of gases under the same conditions of temperature and pressure contain the same number of molecules.

Page 183. Problems.

1. (a) $A = C + 273$
$A = 0 + 273$

$\therefore$ A reading is 273°

(b) 293° A (c) 373° A

(d) 73° A (e) 0° A

2. (a) $C = A - 273$
$C = 100 - 273 = -173$

$\therefore$ Centigrade reading is -173°

(b) $-273^\circ C$ (c) $0^\circ C$ (d) $100^\circ C$

3. $$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

(a) $$\frac{400 \text{ cc}}{293^\circ \text{ A}} = \frac{V_2}{273^\circ \text{ A}}$$

$$V_2 = \frac{400 \text{ cc}}{293^\circ A} \times 273^\circ A = 373 \text{ cc.}$$

or

since the temperature is decreased by $20C^\circ$ the volume is decreased to $\frac{273}{293}$ of 400 cc. = 373 cc.

(b) $$\frac{400 \text{ cc}}{293^\circ \text{ A}} = \frac{V_2}{373^\circ \text{ A}}$$

$$V_2 = \frac{400 \text{ cc}}{293^\circ \text{ A}} \times 373^\circ \text{ A} = 509 \text{ cc.}$$

4. $V_1P_1 = V_2P_2$ $(T = K)$

(a) $100 \text{ cc} \times 760 \text{ mm.} = V_2 \times 700 \text{ mm.}$

$$V_2 = \frac{100 \text{ cc} \times 760 \text{ mm}}{700 \text{ mm}} = 109 \text{ cc.}$$

or

since the pressure is reduced by 60 mm. the volume is increased to $\frac{760}{700} \times 100$ cc = 109 cc.

4. (b) $500 \text{ cc.} \times 740 \text{ mm} = V_2 \times 760 \text{ mm}$

$$V_2 = \frac{500 \text{ cc} \times 740 \text{ mm}}{760 \text{ mm}} = 487 \text{ cc.}$$

(c) $3.50 \text{ l.} \times 700 \text{ mm} = V_2 \times 760 \text{ mm}$

$$V_2 = \frac{350 \text{ l.} \times 700 \text{ mm}}{760 \text{ mm}} = 3.2 \text{ l.}$$

(d) $4.75 \text{ cu.ft.} \times 76 \text{ cm.} = V_2 \times 80 \text{ cm.}$

$$V_2 = \frac{4.75 \text{ cu.ft.} \times 76 \text{ cm.}}{80 \text{ cm.}} = 4.5 \text{ cu.ft.}$$

(e) $10.0 \text{ cu.ft.} \times 30 \frac{\text{lb.wt.}}{\text{sq.in.}} = V_2 \times 40 \frac{\text{lb. wt.}}{\text{sq. in.}}$

$$V_2 = \frac{10.0 \text{ cu.ft.} \times 30 \frac{\text{lb.wt.}}{\text{sq.in.}}}{40 \frac{\text{lb.wt.}}{\text{sq.in.}}} = 7.5 \text{ cu.ft.}$$

(f) $22.4 \text{ l.} \times 1 \text{ at.} = V_2 \times 1.5 \text{ at.}$

$$V_2 = \frac{22.4 \text{ l.} \times 1 \text{ at.}}{1.5 \text{ at.}} = 15 \text{ l.}$$

5. $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ $(P = K)$

(a) $\frac{500 \text{ cc.}}{273^{\circ} \text{ A}} = \frac{V_2}{300^{\circ} \text{ A}}$

$$V_2 = \frac{500 \text{ cc.} \times 300^{\circ}\text{A}}{273^{\circ}\text{A}} = 549 \text{ cc.}$$

(b) $\frac{4.5 \text{ cu.ft.}}{288^{\circ}\text{A}} = \frac{V_2}{373^{\circ}\text{A}}$

$$V_2 = \frac{4.5 \text{ cu.ft.} \times 373^{\circ}\text{A}}{288^{\circ}\text{A}} = 5.8 \text{ cu.ft.}$$

(c) $\frac{22.4 \text{ l.}}{300^{\circ}\text{A}} = \frac{V_2}{273^{\circ} \text{ A}}$

$$V_2 = \frac{22.4 \text{ l.} \times 273^{\circ}\text{A}}{300^{\circ}\text{A}} = 20.4 \text{ l.}$$

(d) $\frac{100 \text{ cc.}}{273^{\circ}\text{A}} = \frac{V_2}{370^{\circ}\text{A}}$

$$V_2 = \frac{100 \text{ cc.} \times 370^{\circ}\text{A}}{273^{\circ} \text{ A}} = 136 \text{ cc.}$$

6. $\frac{V_I P_I}{T_1} = \frac{V_2 P_2}{T_2}$

(a) $\frac{22400 \text{ cc} \times 760 \text{ mm}}{273^{\circ}\text{A}} = \frac{V_2}{300^{\circ}\text{A}} \times 700 \text{ mm.}$

$$V_2 = \frac{22400 \text{ cc} \times 760 \text{ mm} \times 300^{\circ}\text{A}}{273^{\circ}\text{A} \times 700 \text{ mm}} = 26{,}725 \text{ cc.}$$

(b) $\frac{300 \text{ cc}}{273^{\circ}\text{A}} \times 760 \text{ mm} = \frac{V_2}{256^{\circ}\text{ A}} \times 740 \text{ mm.}$

$$V_2 = \frac{300 \text{ cc} \times 760 \text{ mm} \times 256^{\circ}\text{A}}{273^{\circ}\text{A} \times 740 \text{ mm}} = 289\text{cc.}$$

(c) $\frac{22.4 \text{ l}}{393^{\circ}\text{A}} \times 900 \text{ mm.} = \frac{V_2 \times 760 \text{ mm}}{273^{\circ}\text{A}}$

$$V_2 = \frac{22.4 \text{ l} \times 900 \text{ mm} \times 273^{\circ}\text{A}}{393^{\circ}\text{A} \times 760 \text{ mm}} = 18.4 \text{ l.}$$

(d) $\frac{9.75 \text{ cu. ft.} \times 70 \text{ cm}}{236^{\circ}\text{A}} = \frac{V_2 \times 76 \text{ cm}}{273\ ^{\circ}\text{A}}$

$$V_2 = \frac{9.75 \text{ cu. ft.} \times 70 \text{ cm} \times 273^{\circ}\text{A}}{236^{\circ}\text{A} \times 76 \text{ cm}} = 10.4 \text{ cu. ft.}$$

(e) $\frac{27.5 \text{ cu. ft.}}{341^{\circ}\text{A}} \times 45 \frac{\text{lb. wt.}}{\text{sq. in.}} = \frac{V_2}{273^{\circ}\text{A}} \times 15 \frac{\text{lb. wt.}}{\text{sq. in.}}$

$$V_2 = \frac{27.5 \text{ cu. ft.} \times 45 \frac{\text{lb. wt.}}{\text{sq. in}} \times 273^{\circ}\text{A}}{341^{\circ}\text{A} \times 15 \frac{\text{lb. wt.}}{\text{sq. in.}}} = 66 \text{ cu. ft.}$$

7. $\frac{V_1P_1}{T_1} = \frac{V_2P_2}{T_2}$

$$\frac{600 \text{ cc} \times 60 \text{ cm}}{300^{\circ}\text{A}} = \frac{V_2 \times 76 \text{ cm}}{273^{\circ}\text{A}}$$

$$V_2 = \frac{600 \text{ cc} \times 60 \text{ cm} \times 273^{\circ}\text{A}}{300^{\circ}\text{A} \times 76 \text{ cm}} = 432 \text{ cc.}$$

8. $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ $(V = K)$

$$\frac{760 \text{ mm}}{293^{\circ}\text{A}} = \frac{P_2}{153^{\circ}\text{A}}$$

$$P_2 = \frac{760 \text{ mm}}{293^{\circ}\text{A}} \times 153^{\circ}\text{A} = 397 \text{ mm.}$$

9. $\frac{V_1P_1}{T_1} = \frac{V_2P_2}{T_2}$

$$\frac{1000\text{ cc.}}{300^\circ A} \times 15\ \frac{\text{lb. wt.}}{\text{sq. in.}} = \frac{V_2}{693^\circ A} \times 45\ \frac{\text{lb. wt.}}{\text{sq. in.}}$$

$$V_2 = \frac{1000\text{ cc.} \times 15\ \frac{\text{lb wt.}}{\text{sq. in}} \times 693^\circ A}{300^\circ A \times 45\ \frac{\text{lb. wt.}}{\text{sq. in.}}} = 770\text{ cc.}$$

10. $V_1P_1 = V_2P_2$ $(T = 20^\circ C)$

$$V_1 \times 39\ \frac{\text{lb. wt.}}{\text{sq. in.}} = V_2 \times 15\ \frac{\text{lb. wt.}}{\text{sq. in.}}$$

$$V_2 = \frac{V_1 \times 39\ \frac{\text{lb. wt.}}{\text{sq. in.}}}{15\ \frac{\text{lb. wt.}}{\text{sq. in.}}} = 2.6\ V_1$$

$\therefore$ New volume of air is 2.6 times as great as the original volume.

11. $$\frac{V_1P_1}{T_1} = \frac{V_2P_2}{T_2}$$

$$\frac{1\text{ l.} \times 760\text{ mm}}{273^\circ A} = \frac{V_2 \times 700\text{ mm}}{300^\circ A}$$

$$V_2 = \frac{1\text{ l.} \times 760\text{ mm} \times 300^\circ A}{273^\circ A \times 700\text{ mm}} = 1.2\text{ l.}$$

1.2 l. of the gas at $27^\circ C$ and 700 mm. have a mass of 1.43gm.

$$\text{Density} = \frac{1.43\text{ gm}}{1.2\text{ l.}} = 1.2\text{ gm/l.}$$

12. Increase in volume = 48.1 cc. - 35.2 cc. = 12.9 cc.

Increase in temperature = $100^\circ C - 0^\circ C = 100C^\circ$.

Increase of 35.2 cc. of the gas for each $1C^\circ$ rise in temperature = 0.129 cc.

Increase of 1 cc. of the gas for each $1C^\circ$ rise in temperature = $\frac{0.129\text{ cc.}}{35.2}$ = 0.0037 cc.

or

Coef. of volume expansion, $\frac{\Delta V}{V \times \Delta t}$ =

$$\frac{12.9\text{ cc.}}{35.2\text{ cc.} \times 100C^\circ} = 0.0037/C^\circ$$

Page 190. Questions.

1. A calorie is the quantity of heat required to raise the temperature of 1 gram of water 1 centrigrade degree.

2. A British thermal unit is the quantity of heat required to raise the temperature of 1 pound of water 1 Fahrenheit degree.

3. 1 B.t.u. = 252 cal.

4. Thermal capacity is the quantity of heat required to raise the temperature of a body one centrigrade degree (Metric system).

5. Steel is a better conductor of heat than wood, so it removes heat from the warm hand at a greater rate than wood.

6. Water equivalent of a body is the mass of water which has the same thermal capacity as the body.

7. Heat is transferred from the hot body to the cold one until both bodies reach the same temperature.

8. Specific heat of aluminum = 0.22.

Specific heat of lead = 0.03

Aluminum will gain or lose heat

$\frac{0.22}{0.03}$ = 7.3 times as rapidly as lead.

9. Wood is a poor conductor of heat.

10. The thermal capacity of a large body is greater than that of a small body.

11. An insulated vessel used for making heat measurements.

12. See pp. 188 and 189.

Page 191. Problems.

1. $H = m \times \Delta t \times s$

(a) $H = 10 \text{ gm} \times 25C^{\circ} \times 0.11 \text{ cal/gm/}C^{\circ}$

$= 27.5$ cal.

(b) $H = 20 \text{ gm} \times 100C^{\circ} \times 1 \text{ cal/gm/}C^{\circ}$

$= 200$ cal.

(c) $H = 50 \text{ gm} \times 880C^{\circ} \times 0.11 \text{ cal/gm/}C^{\circ}$

$= 4840$ cal.

2. $H = m \times \Delta t \times s$

(a) $H = 15 \text{ lb} \times 110F^{\circ} \times 1 \text{ B.t.u./lb/}F^{\circ}$

$= 1650$ B.t.u.

(b) $H = 20 \text{ lb.} \times 150F^{\circ} \times 0.03 \text{ B.t.u./lb.}F^{\circ}$

$= 90$ B.t.u.

(c) $H = 10 \text{ lb.} \times 80 \times \frac{9}{5} F^{o} \times 0.22 \text{ B.t.u./lb./F}^{o}$

$= 317$ B.t.u.

or

Heat required to raise 1 lb. of aluminum 1 F^{o} = 2.2 B.t.u.

Heat required to raise 10 lb. of aluminum through 1 F^{o} = 2.2 B.t.u.

Heat required to raise 10 lb. of aluminum 144 F^{o} = 317 B.t.u.

3. (a) Heat required to raise 1 gm of water 1 C^{o} is 1 cal.

Heat required to raise 10 gm of water 1 C^{o} is 10 cal.

$\therefore$ thermal capacity of 10 cal. of water is 10 cal.

or

$H = m \times \Delta t \times s$ where $\Delta t = 1\ C^{o}$

$= 10 \text{ gm} \times 1\ C^{o} \times 1 \text{ cal/gm/}C^{o} = 10$ cal.

(b) Thermal capacity $= 150 \text{ gm} \times 1\ C^{o} \times 0.22 \text{ cal/gm/}C^{o}$

$= 33$ cal.

(c) Thermal capacity of 2.5 pounds of iron

$= 2.5 \text{ lb.} \times 1\ F^{o} \times 0.11 \text{ B.t.u./ lb./}F^{o}$

$= 0.275$ B.t.u.

4. Water equivalent = M S (Thermal capacity ratio definition of specific heat is used.)

(a) Water equivalent of 120 gm of copper = m.s.

$= 120 \text{ gm} \times 0.09 = 10.8$ gm.

(b) Water equivalent of 20 gm of glass

$= 20 \text{ gm} \times 0.20 = 4.0$ gm.

(c) Water equivalent of 2 lb. of iron

$= 2 \text{ lb.} \times 0.11 = 0.22$ lb.

5. $H = m \times \Delta t \times s$

$= 40 \text{ gm} \times 60C^{o} \times 0.06 \text{ cal/gm/}C^{o}$

$= 144$ cal.

6. Heat lost by metal = heat gained by water.

$100 \text{ gm } (100^{o}C - 15^{o}C)\ S = 120 \text{ gm } (15^{o}C - 10^{o}C) \times 1 \text{ cal/gm/}C^{o}$

$100 \text{ gm} \times 85C^o \times S = 120 \text{ gm} \times 5C^o \times 1 \text{ cal/gm/}C^o$

$$S = \frac{120 \text{ gm} \times 5C^o \times 1 \text{ cal/gm/}C^o}{100 \text{ gm} \times 85C^o}$$

$$= 0.07 \text{ cal/gm/}C^o$$

7. Heat lost by iron = heat gained by water.
   $89 \text{ gm } (90^oC - 20^oC)\ S = 70 \text{ gm } (20^oC - 10^oC) \times 1 \text{ cal/gm/}C^o$

   $89 \text{ gm} \times 70C^o \times S = 70 \text{ gm} \times 10C^o \times 1 \text{ cal/gm/}C^o$

   $$S = \frac{70 \text{ gm} \times 10C^o \times 1 \text{ cal/gm/}C^o}{89 \text{ gm} \times 70\ C^o} = 0.11 \text{ cal/gm/}C^o$$

8. Heat lost by hot water = heat gained by cold water.
   $100 \text{ gm } (95^oC - t)\ 1 \text{ cal/gm/}C^o = 80 \text{ gm } (t - 20^oC)\ 1 \text{ cal/gm/}C^o$

   $9500^oC - 100\ t = 80\ t - 1600^oC$

   $11100^oC = 180\ t$

   $t = 61.7^oC$

9. Heat required to raise the temperature of calorimeter
   $= 100 \text{ gm} \times 8C^o \times 0.09 \text{ cal/gm/}C^o = 72 \text{ cal.}$
   Heat required to raise the temperature of water
   $= 100 \text{ gm} \times 8C^o \times 1 \text{ cal/gm/}C^o = 800 \text{ cal.}$
   Total heat required = 872 cal.

10. Heat required to raise the temperature of the kettle
    $= 2.5 \text{ lb.} \times 144F^o \times 0.22 \text{ B.t.u./lb./}F^o = 79.2 \text{ B.t.u.}$
    Heat required to raise the temperature of water
    $= 6 \text{ lb.} \times 144F^o \times 1 \text{ B.t.u./lb./}F^o = 864 \text{ B.t.u.}$
    Total heat required = 943.2 B.t.u.

11. Let the specific heat of calorimeter = S
    Heat gained by calorimeter = $200 \text{ gm } (50^oC - 10^oC)\ S$.
    Heat gained by cool water = $50 \text{ gm } (50^oC - 10^oC)\ 1 \text{ cal/gm/}C^o$
    Heat lost by hot water = $100 \text{ gm } (80^oC - 50^oC)\ 1 \text{ cal/gm/}C^o$
    But heat gained = heat lost.
    $(200 \text{ gm} \times 40C^o \times S) + (50 \text{ gm} \times 40C^o \times 1 \text{ cal/gm/}C^o)$
    $= (100 \text{ gm} \times 30C^o \times 1 \text{ cal/gm/}C^o)$

$8000\ S\ gm\ C^o = 2000\ cal + 3000\ cal.$

$$S = \frac{5000\ cal}{8000\ gm\ C^o} = 0.625\ cal/gm/C^o.$$

12. Heat gained by calorimeter = 10 gm ($25^oC - 10^oC$) $1\ cal/gm/C^o$
= 150 cal.
Heat gained by liquid = 80 gm ($25^oC - 10^oC$) S = $1200\ S\ gm\ C^o$
Total heat gained = $1200\ S\ gm\ C^o$ + 150 cal.
Heat lost by metal = 100 gm ($100^oC - 25^oC$) $0.09\ cal/gm/6^o$
= 100 gm x $75C^o$ x 0.09 cal/gm/C^o.
= 675 cal.
150 cal + $1200\ S\ gm\ C^o$ = 675 cal.

$$S = \frac{525\ cal}{1200\ gm\ C^o} = 0.44\ cal/gm/C^o.$$

CHAPTER 21. LATENT HEAT

Page 195. Questions.

1. A physical change of a substance - from a solid to a liquid to a gas or vice versa. Some substances, like CO_2, change directly from a solid state to a gaseous state.

2. No. Ice at 0^o will take in 80 cal/gm of heat without a change in temperature. Water at 100°C will take in 540 calories of heat per gram without a change in temperature.

3. Crystalline substances have definite melting or freezing points. Non-crystalline substances do not.

4. Heat is a form of energy. When a solid is heated, its atoms are energized and their velocities are increased. This increase in velocity causes the attractive forces between the atoms to be overcome and so the atoms move out of their rigid positions.

5. The amount of heat required to change a unit mass of a solid substance to a liquid without a change in temperature.

6. A gram of steam gives off 540 calories of heat in changing to water. Thus, each gram of steam possesses 540 calories of heat more than one gram of water at the same temperature (100°C).

7. Water gives off heat in cooling and freezing.

8. The water vapour present in the air gives off heat in cooling and solidifying.

9. The heat energy builds up the potential energy of the molecules.

10. Evaporation is the change from a liquid to a gaseous state. In order that a liquid change to a gas it must absorb heat. This causes cooling.

11. Latent means "hidden". "Hidden" refers to the fact that although each gram of ice at 0^oC absorbs 80 calories of heat before melting, there is no noticeable change in temperature.

Page 195. Problems.

1. (a) $10 \text{ gm} \times 80 \frac{\text{cal}}{\text{gm}} = 800 \text{ cal.}$

(b) $10 \text{ gm} \times 540 \frac{\text{cal}}{\text{gm}} = 5400 \text{ cal.}$

(c) Specific heat of ice = 0.5 cal/gm/C^o

$H = m \times \Delta t \times s$

$= 20 \text{ gm} \times 20C^o \times 0.5 \text{ cal/gm/}C^o$

$= 200 \text{ cal.}$

(d) Heat required to raise the temperature of 10 gm of ice from -20^oC to 0^oC = 10 gm $(0 - - 20^oC)$ 0.5 cal/gm/C^o

$= 100 \text{ gm} \times 20\ C^o \times 0.5 \text{ cal/gm/}C^o = 100 \text{ cal.}$

Heat required to melt 10 gm of ice at 0^oC to water at 0^oC

$= 10 \text{ gm} \times 80 \text{ cal/gm} = 800 \text{ cal.}$

Heat required to raise the temperature of 10 gm of water from 0^oC to 80^oC = $10 \text{ gm} \times 80\ C^o \times 1 \text{ cal/gm/}C^o = 800 \text{ cal.}$

Total heat = 100 cal + 800 cal + 800 cal = 1700 cal.

(e) Heat required to raise the temperature of 15 gm of ice from -5^oC to 0^oC = $15 \text{ gm} \times 5C^o \times 0.5 \text{ cal/gm/}C^o = 37.5 \text{ cal.}$

Heat required to melt 15 gm of ice at 0^oC to water at 0^oC = $15 \text{ gm} \times 80 \frac{\text{cal}}{\text{gm}} = 1200 \text{ cal.}$

Heat required to raise the temperature of 15 gm of water from $0^{o}C$ to $100^{o}C$ = 15 gm x $100C^{o}$ x 1 cal/gm/C^{o} = 1500 cal.

Heat required to change 15 gm of water at $100^{o}C$ to steam at $100^{o}C$ = 15 gm x 540 cal/gm = 8100 cal.
Total heat required = 10837.5 cal.

(f) Heat required to raise the temperature of 20 gm of ice from $-10^{o}C$ to $0^{o}C$ = 20 gm x $10C^{o}$ x 0.5 cal/gm/C^{o} = 100 cal.

Heat required to melt 20 gm of ice at $0^{o}C$ to water at $0^{o}C$ = 20 gm x 80 $\frac{cal}{gm}$ = 1600 cal.

Heat required to raise the temperature of 20 gm of water from $0^{o}C$ to $100^{o}C$ = 20 gm x $100C^{o}$ x 1 cal/gm/C^{o} = 2000 cal.

Heat required to change the water to steam at $100^{o}C$ = 20 gm x 540 cal/gm = 10800 cal.

Heat required to raise the temperature of 20 gm of steam from $100^{o}C$ to $120^{o}C$ = 20 gm x $20C^{o}$ x 0.5 cal/gm/C^{o} = cal.

Heat required to change 20 gm of ice at $-10^{o}C$ to steam at $120^{o}C$ = 100 cal + 1600 cal + 2000 cal 10800 cal + 200 cal = 14,700 cal.

2. 3.5 lb. x 972 $\frac{B.t.u.}{lb}$ = 3402 B.t.u.

3. Heat required to melt 2 lb. of ice at $32^{o}F$ = 2 lb. x 144 B.t.u./lb. = 288 B.t.u.

   Heat required to raise the temperature of 2 lb. of water from $32^{o}F$ to $212^{o}F$ = 2 lb. x $180F^{o}$ x 1 B.t.u./lb./F^{o} = 360 B.t.u.

   Heat required to change the water to steam = 2 lb. x 972 B.t.u./lb. = 1944 B.t.u.
   Heat required to change 2 lb. of ice at $32^{o}F$ to $212^{o}F$. = 2592 B.t.u.

4. Heat given off in changing 20 gm of steam at $100^{\circ}C$ to water at $100^{\circ}C$ = 20 gm x 540 cal/gm = 10800 cal.

Heat given off in changing 20 gm of water from $100^{\circ}C$ to $0^{\circ}C$ = 20 gm x $100C^{\circ}$ x 1 cal/gm/C° = 2000 cal.

Heat given off in changing 20 gm of water at $0^{\circ}C$ to ice at $0^{\circ}C$ = 20 gm x 80 cal/gm = 1600 cal.

Heat given off in changing 20 gm of steam at $100^{\circ}C$ to ice at $0^{\circ}C$ = 14,400 cal.

5. Let the resulting temperature = t.

Heat given off by steam in condensation = 24 gm x 540 cal/gm = 12,960 cal.

Heat given off by the cooling of water = 24 gm ($100^{\circ}C$ - t) cal/gm/C°. Total heat given off = 12,960 cal + 2400 cal - 24 t cal/C°.

Heat gained by 500 gm of water = 500 gm (t - $20^{\circ}C$) cal/gm/C° = 500 t cal/C° - 10,000 cal.

But heat gained = heat lost.

500 t cal/C° - 10,000 cal = 12,960 cal + 2400 cal - 24 t cal/C°

524 t cal/C° = 25,360 cal.

$$t = \frac{25,360 \text{ cal}}{524 \text{ cal}/C^{\circ}} = 48.4\ C^{\circ} \text{ above 0 or } 48.5^{\circ}C.$$

6. Let the latent heat of vaporization of water = X.

Heat lost by 10 gm of steam at $100^{\circ}C$ in condensing = 10 gm x X = 10X gm.

Heat lost by 10 gm of water in lowering temperature from $100^{\circ}C$ to $40^{\circ}C$ = 10 gm x $60C^{\circ}$ x 1 cal/gm/C° = 600 cal.

Heat gained by 200 gm of water in raising its temperature from $10^{\circ}C$ to $40^{\circ}C$ = 200 gm x $30C^{\circ}$ x 1 cal/gm/C° = 6000 cal.

But heat lost = heat gained.

10X gm + 600 cal = 6000 cal.

6. $10X\ gm = 5400\ cal.$

$X = \frac{5400\ cal}{10\ gm} = 540\ cal/gm$

or

If latent heat of vaporization = X cal/gm, then

$10\ gm \times X\ cal/gm + 600\ cal = 6000\ cal.$

$10\ X\ cal + 600\ cal = 6000\ cal.$

$10\ X = 5400 \quad X = 540$

$\therefore$ latent heat of vaporization = $540\ \frac{cal}{gm}$

7. Let the number of grams of steam = M gm.

Heat lost by steam at 100°C in condensing and having its temperature lowered to 50°C = $M\ gm \times 540\ cal/gm + M\ gm \times 50C^\circ \times 1\ cal/gm/C^\circ = 540\ M\ cal + 50\ M\ cal = 590\ M\ cal.$

Heat gained by 200 gm of water = $200\ gm \times 30C^\circ \times 1\ cal/gm/C^\circ = 6000\ cal.$

$590\ M\ cal = 6000\ cal. \quad M = \frac{6000\ cal}{590\ cal} = 10.2$

$\therefore$ mass of steam = 10.2 gm.

8. Heat lost by steam = $M \times 540\ cal/gm + M \times 100C^\circ \times 1\ cal/gm/C^\circ = 540\ M\ cal/gm + 100\ M\ cal/gm.$

Heat gained by ice = $20\ gm \times 80\ cal/gm = 1600\ cal.$

Heat lost = heat gained.

$640\ M\ cal/gm = 1600\ cal.$

$M = \frac{1600\ cal}{640\ cal/gm} \qquad M = 2.5\ gm$

9. 100 gm of water in cooling from 50°C to 0°C give off $100\ gm \times 50C^\circ \times 1\ cal/gm/C^\circ = 5000\ cal$ of heat.

5000 cal of heat will melt $\frac{5000\ cal}{80\ cal/gm}$ or 62.5 gm of ice.

Since there is not enough heat to melt all the ice, the resulting temperature is 0°C.

A mixture of (200 gm - 62.5 gm) -- 137.5 gm of ice and (100 gm + 62.5 gm) 162.5 gm of water at 0°C is the result.

10. Let mass of water = M gm.

To melt all the ice the least possible temperature is 0°C

Heat lost = heat gained.

$M\ gm \times 100C^\circ \times 1\ cal/gm/C^\circ = 100\ gm \times 80\ cal/gm$

$100\ M\ cal = 8000\ cal. \quad M = 80$

$\therefore$ Mass of water required = 80 gm.

11. Let the heat of fusion of ice = X cal/gm

Heat lost by calorimeter = 200 gm x 40C° x 0.10 cal/gm/C° = 800 cal.

Heat lost by 400 gm of water = 400 gm x 40C° x 1 cal/gm/C° = 16000 cal.

Total heat lost = 16,800 cal.

Heat gained by ice in melting = 200 gm x X cal/gm = 200 X cal.

Heat gained by water formed from ice
= 200 gm x 5C° x 1 cal/gm/C° = 1000 cal.

Heat gained = heat lost.
200 X cal + 1000 cal = 16,800 cal.
200 X = 15,800 X = 79.

$\therefore$ Heat of fusion of ice = 79 cal/gm

12. Let the latent heat of vaporization of water = X° cal/gm

Heat lost:

(1) by steam in condensing = 23.5 gm x X cal/gm = 23.5 X cal.

(2) by water formed from steam = 23.5 gm x 60C° x 1 cal/gm/C° = 1410 cal.

Total heat lost = 23.5 X cal + 1410 cal.

Heat gained:

(1) by calorimeter = 120 gm x 34C° x 0.1 cal/gm/C° = 408 cal.

(2) by 402 gm of water = 402 gm x 34C° x 1 cal/gm/C° = 13,668 cal.

Total heat gained = 14,076 cal.

But heat lost = heat gained.

(23.5 X + 1410) cal = 14,076 cal.

23.5 X = 12666 $X = \frac{12666}{23.5} = 539$

$\therefore$ the latent heat of vaporization of water = 539 cal/gm.

UNIT 7: SOUND

CHAPTER 23. PRODUCTION, PROPAGATION, AND VELOCITY OF SOUND

Page 205. Questions.

1. Sound waves are produced by vibrating matter.
2. A material medium - solid, liquid or gas.
3. Light travels much faster than sound.
4. The ground acts as a better medium for transmitting sound than air.
5. Light travels much faster than sound.

6. If a steep bank or a large hill exists on the opposite side of the lake, one could shout and find the time it takes for his echo to be heard. The time of the echo divided by two and multiplied by the velocity of sound will give the distance across the lake.

Or, one could watch a train whistle blow on the opposite side of the bank. Find the time difference between the sighting of the steam and the hearing of the sound and multiply this time by the velocity of sound.

7. Some of the factors which make hearing a speaker more difficult outside than inside a building are:

(1) Outside, the sound travels in all directions, while inside, the sound is directed at the listener by the walls.
(2) Interference by other sound waves.
(3) Air currents carry sound.

8. A megaphone directs sound waves in one direction.

9. To prevent the sound waves from reflecting off smooth walls and thus forming echoes.

10. Same as heavy drapes in question 9. The audience absorbs sound waves and prevents the formation of echoes.

Page 205. Problems.

1. Increase in speed of sound is approximately 2 ft/sec for each centigrade degree rise in temperature. Speed of sound at $0^{\circ}C$ = 1090 ft/sec.

(a) 1090 ft/sec + $10C^{\circ}$ x 2 ft/sec/C° = 1110 ft/sec.
(b) 1150 ft/sec.
(c) 1050 ft/sec.
(d) 1130 ft/sec.

2. Increase in speed of sound is approximately 0.6 meters/sec/C°.

Speed of sound at $0^{\circ}C$ is 332 meters/sec.

(a) 332 m/sec + $20C^{\circ}$ x 0.6 m/sec/C° = 344 m/sec.
(b) 356 m/sec.
(c) 326 m/sec.

3. Time for sound to travel one way = 1.5 sec.

Velocity of sound at $68^{\circ}F$ ($20^{\circ}C$) = 1130 ft/sec.

Distance = 1.5 sec x 1130 ft/sec = 1695 ft.

4. Time = 5 seconds

Velocity = 1130 ft/sec.

Distance = 5 sec x 1130 ft/sec = 5650 ft.

5. Time for sound to reach bottom of ocean is 2 sec.

Speed of sound in water = 4800 ft/sec.

Depth of ocean = 2 sec x 4800 ft/sec = 9600 ft.

6. Time = $\frac{1}{2}$ sec.

Velocity of sound = 1100 ft/sec.

Depth of water level = $\frac{1}{2}$ sec. x 1100 ft/sec. = 550 ft.

7. Time for sound to reach cliff = 3 sec.

Distance of cliff = 3300 ft.

$$\text{Velocity} = \frac{\text{distance}}{\text{time}}$$

$$\text{Velocity of sound} = \frac{3300 \text{ ft.}}{3 \text{ sec.}} = 1100 \text{ ft/sec.}$$

CHAPTER 24. TRANSMISSION OF SOUND

Page 209. Questions.

1. Longitudinal waves.

2. In transverse vibrations the particles vibrate at right angles to the path along which the wave travels, while in longitudinal vibrations motion is along the path of wave.

3. (a) Wave length is the distance between two consecutive particles in the same phase.

(b) Amplitude is the maximum displacement from the mean position.

(c) Frequency is the number of complete vibrations or cycles per second.

(d) Phase refers to particles at the same distance from their mean positions and moving in the same direction.

(e) Period is the time required for one cycle.

(f) Condensations are the compressed regions in longitudinal waves.

Rarefactions are the stretched regions.

(g) Standing waves do not seem to move.

(h) Node is the point of zero motion.

Antinode is the region of maximum motion.

(i) Fundamental tone is the tone of lowest frequency emitted when the string vibrates in one segment.

Overtones are formed by frequencies which are multiples of the fundamental.

(j) Beats are recurrences of maximum loudness.

4. From everyday experiences. High pitch and low pitch sounds produced the same distance away are heard at the same time.

5. V = nl.

6. Yes. We know that the frequency is constant, yet as sound travels faster in warm air the wave length must be longer.

7. By eliminating the beats.

Page 210. Problems.

1. $l = 2$ in; $n = 6/\text{sec}$.

   $v = nl = \frac{6}{\text{sec}} \times 2 \text{ in.} = 12 \text{ in/sec}$.

2. $n = 320$ sec. $l = 3.56$ ft.

   (a) $V = nl = 320/\text{sec} \times 3.56 \text{ ft} = 1139 \text{ ft/sec}$.

   (b) Velocity of sound at $0^{\circ}C = 1090$ ft/sec.

   Increase in velocity = 49 ft/sec.

   Temperature above $0^{\circ}C = \frac{49 \text{ ft/sec}}{2 \text{ ft/sec/C}^{\circ}} = 24.5C^{\circ}$

   $\therefore$ temperature = $24.5^{\circ}C$

3. $n = 275/\text{sec}$; $t = 5^{\circ}C$

   $v = 1100$ ft/sec.

   $l = \frac{v}{n} = \frac{1100 \text{ ft/sec}}{275/\text{sec}} = 4$ ft.

4. $n = 256/\text{sec}$. $v = 1150$ ft/sec.

   $l = \frac{1150 \text{ ft/sec}}{256 \text{ sec}} = 4.5$ ft.

5. $n = 264$ vibrations/sec.

   $l = 4.2$ ft/vibration.

   $v = nl = 264 \text{ vib/sec} \times 4.2 \text{ ft/vib} = 1109 \text{ ft/sec}$.

6. $l = 175$ cm. $t = 30^{\circ}C$. $n = ?$

   $v = 332 \text{ m/sec} + 0.6 \times 30 \text{ m/sec} = 350 \text{ m/sec}$.

   $n = \frac{v}{l} = \frac{35000 \text{ cm/sec}}{175 \text{ cm}} = 200/\text{sec}$.

7. 4 segments = 2 wave lengths.

   2 wave lengths = 24 ft.

   $l = 12$ ft. $n = 2/\text{sec}$.

   $v = 12 \text{ ft} \times 2/\text{sec} = 24 \text{ ft/sec}$.

8. Frequency of beats = 264/sec - 256/sec = 8/sec.

9. Frequency of second tuning fork = 440/sec - 5/sec. = 435/sec.

10. Frequency of first string = 512/sec + 4/sec = 516/sec.

CHAPTER 25. DISTINGUISHING FEATURES OF SOUND

Page 213. Questions.

1. Pitch is determined by the frequency of vibration.

2. Amplitude of the waves and the area of the vibrating body.

3. Loudness is determined by intensity, distance and type of medium.

4. Quality, which in turn depends upon the number of overtones present in the sound wave.

5. As the train is approaching, the sound waves striking the listener are shortened proportionally to the speed of the train. The listener receives more vibrations per second.

6. As the train moves away the sound waves are lengthened and fewer vibrations per second reach the target.

7. A dog can hear and interpret sounds of high frequency which are too high to be audible to a human.

8. 20 cycles/sec to 20,000 cycles/sec.

9. Speed of the saw varies with varying loads.

10. Sound waves are given off in all directions from a vibrating body, spreading over greater areas as they move farther away from the source. A megaphone concentrates the waves in one direction.

11. The surface of the piano vibrates in resonance with the vibrating strings. These vibrations are interpreted by touch.

12. A child's vocal chords are underdeveloped. They are smaller and therefore vibrate at a greater rate.

CHAPTER 26. RESONANCE: SYMPATHETIC VIBRATIONS

Page 217. Questions.

1. One body sets another into vibration. This occurs when the period of the two is the same or when one is a multiple of the other.

2. The note of the singer has the same period as the violin string, or one is a multiple of the other. The condensations and rarefactions of air molecules produced by the singer are in resonance with the vibrations of the string causing the string to vibrate.

3. The note produced by the lowest frequency of the vibrating body.

4. Frequencies which are multiples of the fundamental.

5. The frequency of the first overtone in a closed tube is 3 times that of its fundamental; in an open tube it is 2 times that of the fundamental.

6. The frequency of the fourth overtone is
(i) 9 times that of the fundamental in a closed tube.
(ii) 5 times that of the fundamental in an open tube.

7. The frequency of the fifth overtone is

(i) 11 times that of the fundamental in a closed tube.
(ii) 6 times that of the fundamental in an open tube.

8. Resonance occurs when the period of vibration of two bodies is the same, or one is a multiple of the other.

9. Lw = 4 lr

10. Lw = 2 lr

11. As the vessel is filled, the resonance tube becomes shorter and so do the sympathetic wave lengths causing the pitch to rise.

12. See page 215. By using a tuning fork of known frequency and finding the length of the air column in resonance. If the tube is closed, use Lw = 4 lr. If open use Lw = 2 lr.

Page 217. Problems.

1. Lw = 4 lr

   2 ft. = 4 lr

   lr = 0.5 ft.

2. Lw = 2 lr

   Lw = 2 x 4 ft = 8 ft.

3. Wave length $= \frac{1152 \text{ ft/sec}}{384/\text{sec}} = 3$ ft.

   Lw = 2 lr

   3 ft = 2 lr

   lr = 1.5 ft.

4. lr = 14 in.   v = 1120 ft/sec

   Lw = 2 lr

   Lw = 2 x 14 in. = 28 in. or 2 1/3 ft.

   $n = \frac{v}{l}$

   $n = \frac{1120 \text{ ft/sec}}{2\ 1/3 \text{ ft.}} = 480/\text{sec.}$

5. lr = 4 ft.

   Lw = 4 lr. = 4 x 4 ft = 16 ft.

6. lr = 12 ft.

   Lw = 2 lr = 2 x 12 ft. = 24 ft.

7. n = 1760/sec.

   Velocity at 25^{o}C = 1140 ft/sec.

   Wave length $= \frac{1140 \text{ ft/sec}}{1760/\text{sec.}} = 0.65$ ft.

   Length of pipe $= \frac{0.65 \text{ ft.}}{2} = 0.33$ ft.

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8. Assuming that the organ pipe is open, the frequency of the first overtone is 512 vib/sec., and the second overtone is 768 vib/sec.

9. lr = 1 ft. n = 256 vib/sec.

Lw = 4 ft.

(a) Velocity of sound = 256 vib/sec x 4 ft/vib = 1024 ft/sec.

(b) Velocity of sound at $0^{\circ}C$ is 1090 ft/sec.

Decrease in velocity = 66 ft/sec.

Decrease in temperature = $33C^{\circ}$

$\therefore$ temperature = $-33^{\circ}C$

10. lr = 1.5 ft.

Lw = 3 ft.

n = 384/sec.

(a) Velocity of sound = 384/sec x 3 ft = 1152 ft/sec.

(b) Velocity of sound at $0^{\circ}C$ = 1090 ft/sec

Increase in velocity = 62 ft/sec.

Increase in temperature = $31C^{\circ}$

$\therefore$ temperature = $31^{\circ}C$

CHAPTER 27. LAWS OF STRETCHED STRINGS

Page 221. Questions.

1. The strings of a piano have different lengths, different tensions, different diameters and are made of materials of different densities.

2. Law of tensions.

3. Law of lengths. Lengths are varied by using fingers.

4. The motion of the bridge changes the length of the string.

5. A thin wire is wrapped around the gut string.

6. In playing a violin or a guitar, the pitch is raised by shortening the vibrating section of the string by using fingers.

Page 222. Problems.

1. The frequency of a string varies inversely as its length. Since the second string is 3/4 as long as the first string, its frequency is 4/3 as great or 400/sec x $\frac{4}{3}$ = 533/sec.

or

$$\frac{n_1}{n_2} = \frac{l_2}{l_1}$$

$$\frac{400/\text{sec}}{n_2} = \frac{3/4\ 1}{1}$$

$$\frac{3}{4} n_2 = 400/\text{sec}$$

$$n_2 = 400/\text{sec} \times \frac{4}{3} = 533/\text{sec.}$$

2. $$\frac{n_1}{n_2} = \frac{l_2}{l_1}$$

$$\frac{200 \text{ vib/sec}}{n_2} = \frac{80 \text{ cm}}{100 \text{ cm}}$$

$$n_2 = 200 \text{ vib/sec} \times \frac{100}{80} = 250 \text{ vib/sec.}$$

3. $$\frac{n_1}{n_2} = \frac{\sqrt{t_1}}{\sqrt{t_2}}$$

$$\frac{512/\text{sec}}{n_2} = \frac{\sqrt{16 \text{ lb.wt.}}}{\sqrt{25 \text{ lb.wt.}}}$$

$$n_2 = 512/\text{sec} \times \frac{5}{4} = 640/\text{sec.}$$

4. $$\frac{n_1}{n_2} = \frac{\sqrt{t_1}}{\sqrt{t_2}}$$

$$\frac{256 \text{ vib/sec}}{320 \text{ vib/sec}} = \frac{\sqrt{25 \text{ lb.wt.}}}{\sqrt{t}}$$

$$\sqrt{t} = \sqrt{25 \text{ lb.wt.}} \times \frac{320}{256}$$

$$\sqrt{t} = \sqrt{25 \text{ lb.wt.}} \times \frac{5}{4}$$

$$\sqrt{t} = \sqrt{25 \text{ lb.wt.}} \times \frac{25}{16} = 39 \text{ lb.wt.}$$

5. $$\frac{n_1}{n_2} = \frac{d_2}{d_1}$$

$$\frac{420/\text{sec}}{n_2} = \frac{0.2 \text{ mm}}{0.8 \text{ mm}}$$

$$n_2 = 420/\text{sec} \times \frac{0.8}{0.2} = 1680/\text{sec.}$$

6. $$\frac{n_1}{n_2} = \frac{\sqrt{D_2}}{\sqrt{D_1}}$$

$$\frac{440/\text{sec}}{n_2} = \frac{\sqrt{10}}{\sqrt{2.5}} = \sqrt{\frac{10}{2.5}} = \sqrt{4} = 2$$

$$n_2 = 220/\text{sec.}$$

7. $\frac{N\ nylon}{N\ steel} = \frac{\sqrt{D\ steel}}{\sqrt{D\ nylon}}$

$\frac{100/sec}{N\ steel} = \sqrt{\frac{7.7}{1.1}} = \sqrt{7} = 2.65$

$N_{steel} = \frac{100/sec}{2.65} = 38/sec.$

8. Frequency of first string = 320 vib/sec. Because the length of the second string is 4 times the first, its frequency is 1/4 as great or 80 vib/sec. Because the diameter of the second string is ½ as great as the first, its frequency is 2 times as great or 160 vib/sec.

or

$$\frac{n_1}{n_2} = \frac{l_2}{l_1} \times \frac{d_2}{d_1}$$

$$\frac{320\ vib/sec}{n_2} = \frac{100\ cm}{25\ cm} \times \frac{0.25\ mm}{0.5\ mm}$$

$$\frac{320\ vib/sec}{n_2} = \frac{2}{1}$$

$$n_2 = 160\ vib/sec.$$

9. $l_1 = 16$ in. $\quad l_2 = 12$ in.

$t_1 = 49$ lb. wt. $\quad t_2 = 100$ lb.wt.

$n_1 = 420$ vib/sec. $n_2 = ?$

$$\frac{n_1}{n_2} = \frac{l_2}{l_1} \times \frac{\sqrt{t_1}}{\sqrt{t_2}}$$

$$\frac{420}{n_2} = \frac{12}{16} \times \sqrt{\frac{49}{100}}$$

$$\frac{420}{n_2} = \frac{12}{16} \times \frac{7}{10} = \frac{21}{40}$$

$$n_2 = 420 \times \frac{40}{21} = 800$$

the frequency of the second string is 800 vibrations/sec.

10. Because the length of the second string is twice the first, its frequency is ½ as great as that of the first. Because its tension is 16 times as great, its frequency is $\sqrt{16}$ or 4 times as great.

Combining the two we find that the frequency of the second is (½ x 4) or 2 times as great as the first.

or

$$\frac{n_1}{n_2} = \frac{l_2}{l_1} \times \frac{\sqrt{t_1}}{\sqrt{t_2}}$$

10. $\dfrac{n_1}{n_2} = \dfrac{2\ 1}{1} \times \dfrac{\sqrt{t}}{\sqrt{16t}}$

$\dfrac{n_1}{n_2} = \dfrac{2\ 1}{1} \times \sqrt{\dfrac{t}{16t}}$

$\dfrac{n_1}{n_2} = 2 \times 1/4 = \frac{1}{2}$ or $n_2 = 2n_1$

11. $\dfrac{N_1}{N_2} = \dfrac{\sqrt{D_2}}{\sqrt{D_1}}$

$\dfrac{120 \text{ vib/sec}}{N_2} = \dfrac{\sqrt{2\ 1/4\ D_1}}{\sqrt{D_1}} = \sqrt{\dfrac{9}{4}} = \dfrac{3}{2}$

$N_2 = 120 \text{ vib/sec} \times \dfrac{2}{3} = 80 \text{ vib/sec}$

12. $\dfrac{n_1}{n_2} = \dfrac{l_2}{l_1} \times \dfrac{\sqrt{t_1}}{\sqrt{t_2}}$

$\dfrac{320 \text{ vib/sec}}{n_2} = \dfrac{80 \text{ cm}}{100 \text{ cm}} \times \dfrac{\sqrt{16 \text{ lb.wt.}}}{\sqrt{9 \text{ lb.wt.}}}$

$\dfrac{320 \text{ vib/sec}}{n_2} = \dfrac{4}{5} \times \dfrac{4}{3} = \dfrac{16}{15}$

$n_2 = 320 \text{ vib/sec} \times \dfrac{15}{16} = 300 \text{ vib/sec.}$

13. Original frequency = 100 vib/sec.

Since the length is multiplied by 4, the frequency becomes 1/4 x 100 vib/sec. = 25 vib/sec.

Since the tension is multiplied by 4, the frequency becomes 25 vib/sec x $\sqrt{4}$ = 50 vib/sec.

Since the diameter is multiplied by 4, the frequency becomes 50 vib/sec x 1/4 = 12.5 vib/sec.

Since the density is multiplied by 4, the frequency becomes 12.5 vib/sec. x $\dfrac{1}{\sqrt{4}}$ = 6.25 vib/sec.

or

$\dfrac{n_1}{n_2} = \dfrac{l_2}{l_1} \times \dfrac{\sqrt{t_1}}{\sqrt{t_2}} \times \dfrac{d_2}{d_1} \times \dfrac{\sqrt{D_2}}{\sqrt{D_1}}$

$\dfrac{100}{n_2} = \dfrac{4\ 1}{1} \times \dfrac{\sqrt{1}}{\sqrt{4\ 1}} \times \dfrac{4d}{d} \times \dfrac{\sqrt{4D}}{\sqrt{D}}$

$\dfrac{100}{n_2} = 4 \times \frac{1}{2} \times 4 \times 2 = 16$

$n_2 = 6.25$

$\therefore$ Frequency is 6.25 vib/sec.

SECTION IV: SUPPLEMENTARY EXERCISES

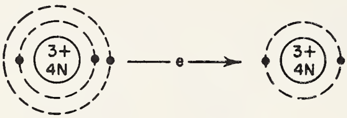
CHEMISTRY

1. (a) Draw diagrams to show how an atom of lithium changes to an ion.

(b) Show this change by an electron transfer equation.

Answer:

(a)



(b) $Li^{o} - e \longrightarrow Li^{+}$

2. (a) Draw diagrams and (b) write electron transfer equations to show how each of the following atoms change to ions:

beryllium	oxygen
sodium	nitrogen
magnesium	fluorine
potassium	chlorine
calcium	

(c) Classify each of the above elements as metals or non-metals. State your reasons.

3. Write electron transfer equations to show how each of the following ions may be discharged to become atoms:

(a) sodium	(e) fluorine
(b) calcium	(f) oxygen
(c) magnesium	(g) sulfur
(d) aluminum	(h) nitrogen

4. (a) Draw diagrams representing atoms of helium, neon and argon.

(b) Why are these gases considered to be inert to chemical reactions?

5. Draw diagrams representing molecules of the following diatomic gases:

(a) hydrogen (b) oxygen (c) fluorine

Why are they called diatomic?

6. Draw diagrams showing how each of the following pairs of elements combine to form molecules. Name the products:

(a) lithium and chlorine
(b) sodium and oxygen
(c) magnesium and oxygen

(d) aluminum and oxygen
(e) carbon and oxygen
(f) sodium and chlorine
(g) calcium and fluorine

7. Why are lithium, sodium and potassium located in group 1A of the periodic chart?

8. Why are sodium, magnesium and aluminum located in period 3 of the periodic chart?

9. In the periodic table, calcium is listed with the following information:

2 8 8 2	20 Ca 40.08

What does each number represent?

10. An element "R" has an atomic number of 12.

(a) In what group is it located?

(b) In what period is it located?

(c) Is it a metal or a non-metal?

11. An element "R" has an atomic number 17.

(a) In what group is it located?

(b) In what period is it located?

(c) Is it metallic or non-metallic?

12. An element "R" is located in group 2 and period 4.

(a) What is its atomic number?

(b) What is its valence?

(c) Is it a metal or non-metal?

(d) Write the formula for its hydride.

(e) Write the formula for its oxide.

13. Elements are divided into two groups A and B as shown in the periodic table. What is the essential difference between group A and group B?

14. The periodic table is known as a chemist's tool. State three uses of the periodic table.

15. Write balanced equations for the following chemical reactions:
Classify each reaction.

(a) Heating mercuric oxide.

(b) Heating a mixture of potassium chlorate and manganese dioxide.

(c) Electrolysis of water.

(d) Heating of iron and sulfur.

(e) Heating of copper and sulfur.

(f) Exploding hydrogen and oxygen in a audiometer.

(g) Action of potassium metal on water.

(h) Action of sodium metal on water.

(i) Action of calcium metal on water.

(j) Action of dilute hydrochloric acid on magnesium.

(k) Action of dilute hydrochloric acid on zinc.

(l) Action of dilute hydrochloric acid on iron.

(m) Action of dilute sulfuric acid on (a) magnesium, (b) zinc, (c) iron.

(n) An iron nail is dropped into a solution of copper sulfate.

(o) A piece of sodium metal is dropped into a solution of ferrous chloride.

(p) Chlorine gas is bubbled through a solution of potassium iodide.

(q) Chlorine gas is bubbled through a solution of potassium bromide.

(r) A solution of silver nitrate is poured into a solution of sodium chloride.

(s) A solution of barium chloride is poured into a solution of sodium sulfate.

(t) Sodium oxide reacts with water.

(u) Calcium oxide reacts with water.

(v) Sulfur trioxide reacts with water.

(w) Carbon dioxide reacts with water.

The above reactions are grouped in groups of 2 or 3. Examine your equations. Group them. What does each group illustrate.

16. (a) Define an acid.

(b) Write ionic equations showing the dissociation of the following acids:

(i) hydrochloric
(ii) sulfuric
(iii) nitric
(iv) acetic
(v) phosphoric

17. (a) Define a base.

(b) Write ionic equations showing the dissociation of following bases:

(i) sodium hydroxide
(ii) potassium hydroxide

(iii) calcium hydroxide
(iv) ferrous hydroxide
(v) aluminum hydroxide

18. (a) Define a salt.

(b) Write ionic equations showing the dissociation of the following salts in water solutions:

(i) sodium chloride
(ii) potassium iodide
(iii) sodium sulfate
(iv) potassium phosphate
(v) ammonium bisulfate

19. Find the percentage composition of each of the following compounds:

(a) water (H = 11.1%; O = 88.9%)

(b) calcium carbonate (Ca = 40%; C = 12%; O = 48%)

(c) sulfuric acid (H = 2.0%; S = 32.7%; O = 65.3%)

(d) sodium hydroxide (Na = 57.5%; O = 40%; H = 2.5%)

(e) zinc phosphate (Zn = 50.5%; P = 16.1%; O = 33.3%)

20. Find the percentage of sulfur in calcium sulfate (23.5)

21. Find the percentage of iron in hematite, Fe_2O_3. (70.0)

22. Find the percentage of water of crystallization in each of the following:

(a) Bluestone, $CuSO_4 \cdot 5H_2O$ (36.1).

(b) Hypo crystals, $Na_2S_2O_3 \cdot 5H_2O$ (36.6)

(c) Washing soda, $Na_2CO_3 \cdot 10H_2O$ (62.9)

(d) Epsom salts, $MgSO_4 \cdot 7H_2O$ (51.2)

23. You bought 100 gm. of washing soda crystals. How many grams of water did you buy? (62.9)

24. 5.0 gm of bluestone crystals are heated until completely dehydrated. What is the weight of the dehydrated copper sulfate? (3.2 gm)

25. Which would be more economical to buy, 10 pounds of washing soda crystals at 5 cents per pound or 5 pounds of anhydrous sodium carbonate at 10 cents per pound? (Anhydrous. You get 1.3 lb. of Na_2CO_3 more for 50 cents).

26. Assuming that the ore bauxite contains 80% pure aluminum oxide, how many tons of aluminum can be extracted from 10 tons of bauxite? (4.2)

27. 499 grams of bluestone crystals are heated to form 319 grams of anhydrous copper sulfate. Find the formula for bluestone crystals. ($CuSO_4 \cdot 5H_2O$)

28. What weight of oxygen can be prepared by the complete decomposition of 10 grams of potassium chlorate? (3.9 gm).

29. What weight of oxygen must combine with 5.0 gm of sulfur to change all the sulfur to sulfur dioxide? (5 gm)

30. What weight of hydrogen must combine with oxygen to form 3.0 grams of water? (0.33 gm)

31. (a) How many grams of water are formed when 20 grams of hydrogen are completely oxidized? (180)

(b) What volume of water is formed? (180 cc)

32. What weight of carbon is required to extract 1 ton of iron from hematite, Fe_2O_3.

$$Fe_2O_3 + 3C \longrightarrow 2Fe + 3CO. \qquad (0.6 \text{ ton})$$

33. Calculate the number of tons of hydrogen needed to produce 10 tons of ammonia. (1.8)

34. Ten grams of hydrogen were produced in the electrolysis of water. How many grams of oxygen were produced in the process? (80)

35. When 10 grams of washing soda crystals were heated, 3.70 gm. of anhydrous sodium carbonate remained. Find the number of molecules of water of crystallization in a molecule of the crystals. (10)

36. Five grams of potassium chlorate are completely decomposed. Find the weight of potassium chloride remaining. (3.0 gm)

37. Find the maximum number of molecules of oxygen which can be produced from 6 molecules of potassium chlorate. (9)

38. A solution of hydrochloric acid is 10% HCl by weight. How many grams of such a solution are required to neutralize 100 gm. of pure potassium hydroxide? (365 gm.)

PHYSICS

MECHANICS OF FLUIDS

1. A rectangular aluminum block 10 cm. x 5 cm. x 2 cm. has a mass of 270 gm. Find the different pressures which it can exert when resting on a horizontal surface. (5.4, 13.5, and 75 gm. wt./cm.^{2})

2. On what area must a force of 10 lb. wt. act to produce a pressure of 50 lb. wt./in^2? (0.2 in.^{2})

3. A rectangular block of iron 6 in. x 2 in. x 3 in. has a mass of 10 lb. Calculate the greatest and the least pressures it can exert when resting on a horizontal surface. (1.67 and 0.55 lb. wt./in.^{2})

4. A water manometer reads 12.4 cm. when connected to a gas main. Find the pressure of the gas in excess of the atmospheric pressure. (12.4 gm.wt./cm.^{2})

5. Express a pressure of 4.2 in. of mercury in lb. wt./in.^{2}. (2.06 lb. wt./in.^{2})

6. Calculate the pressure in tons wt./ft.^{2} at the bottom of a dam when the water head is 36 ft. (1.0 ton wt/ft.^{2})

7. Find the velocity ratio of a hydraulic lift which is fitted with cylinders of diameter 2 in. and 24 in. (144)

8. A hydraulic lift having a piston of 30 cm. diameter is operated by a piston of 6 cm. diameter. What resistance can be overcome if a force of 3 Kg. wt. is applied to the small piston? (Neglect friction and weight of machine.) (75 Kg. wt.)

9. Calculate the pressure of sea water at a depth of 100 ft. below the surface. Density of sea water is 60 lb./ft.^{3} (44.4 lb. wt/in.^{2})

10. The diameters of a hydraulic press are 2 in. and 8 in. respectively. A force of 20 lb. wt. is applied to the smaller piston. Calculate:

(a) the pressure in the press (6.37 lb. wt./in.^{2})

(b) the force exerted on the larger piston. (320 lb.wt.)

11. A cube of steel 2 cm. along each edge has a mass of 64 gm. Find:

(a) its density. (8 gm/cm^3)

(b) its specific gravity. (8)

12. Find the pressure on a water faucet if the water head is 72 ft. (31.2 lb.wt./in.^{2})

13. A dam 20 ft. wide holds 10 ft. of water. Find the thrust on the dam. (62,400 lb. wt.)

14. In a hydraulic press the area of the large piston is 10 in.^{2} and the small piston is 0.5 in.^{2}. What resistance can be overcome is a force of 20 lb. wt. is applied to the small piston? (400 lb.wt.)

15. Express a barometric pressure of 30 in. of mercury in lb. wt. in.^{2}. (14.8)

16. Express a barometric pressure of 30 in. of mercury in ft. of water. (34)

17. If the average density of air is 0.001293 gm/cm.^{3} find the approximate height of the atmosphere if the barometer reads 76 cm. of mercury. (8.0 Km.)

18. A tube closed at one end and open at the other end was filled with mercury and inverted into a beaker of mercury. The mercury stood at 74 cm. above the level of mercury in the beaker On depressing the tube into the beaker so that the volume of the space above the mercury is one-third its original value, the mercury level is only 71 cm. above that in the beaker. Find the atmospheric pressure. (75.5 cm.)

19. The atmospheric pressure on a mountain top is 50 cm. of mercury. What is the maximum height to which water can be pumped up? (680 cm.)

20. If the average density of air is 0.0012 gm/cm^3, how high must one climb to realize a drop of 1 in. of mercury in atmospheric pressure? (944 ft.)

21. Express a resource of 76 cm. of mercury in gm.wt./in.^{2} (1033).

22. An aneroid barometer has an effective area of 30 cm.^{2}. By how much will the force exerted upon it by the atmosphere change when the mercury in a barometer falls one cm.? (408 gm.wt.)

23. Find the barometric height in a water barometer when a mercury barometer reads 75.0 cm. (9.9 m.)

24. A lake is 6.8 meters deep. Calculate the pressure at the bottom of this lake when the atmospheric pressure reads 74 cm. of mercury. (1686 gm. wt./cm.^{2})

25. The density of sulfur dioxide gas is 2.86 gm./litre when its pressure is 76 cm. of mercury. Find its density when the pressure is 114 cm. of mercury. (4.29 gm./l.)

HEAT

26. If the distance between the fixed points on a thermometer is 15 cm., how high above the lower fixed point will the mercury stand at 20°C.? (3 cm.)

27. A gas occupies 30 cm.^{3} at a temperature of 27°C. What is its volume at -10°C. if the pressure remains constant? (26.3 cm.^{3})

28. Find the amount of heat liberated by 90 gm. of mercury in cooling from 220°C. to 20°C. (Specific heat of mercury is 1/30.) (600 cal.)

29. The coefficient of linear expansion of aluminum is 0.000023/C°. What is the coefficient of volume expansion? (0.000069/C°)

30. If 20 cm.^{3} of water at 80°C. are mixed with 100 cm.^{3} of water at 20°C., find the final temperature. (30°C.).

31. The volume of a given mass of gas is 1000 cm.^{3} at 20°C. At what temperature will its volume be doubled? (313°C.).

32. How much heat is released when 100 gm. of iron are cooled from 100°C. to 20°C.? Specific heat of iron is 0.113. (904 cal.)

33. If the pressure of a given mass of gas is multiplied by 4, temperature remaining constant, what is the change in its volume? (1/4)

34. Change 68°F. to degrees absolute. (293°A)

35. The coefficient of linear expansion of steel is 0.000011/C°. A steel tape is 100 ft. long at 0°C. What will be the increase in length when the temperature is 25°C. (0.0275 ft.)

36. What is the amount of heat required to change 10 gm. of ice at 0°C. to steam? (7200 cal.)

37. 12 gm. of steam are blown into a large block of ice. Neglecting heat losses find the maximum number of grams of ice that could be melted. (96)

38. Find the water equivalent of a glass stirring rod having a mass of 20 gm. (3.2 gm.)

39. What means of water at 80°C. will just melt 80 gm. of ice? (80 gm.)

40. In using Loy's apparatus it was found that the volume of a given mass of gas was 35.2 cm.^{3}, at 0°C. and 48.1 cm.^{3} at 100°C. Find the mean coefficient of expansion of the gas between 0°C. and 100°C. (0.0037)

41. What is the number of B.t.u.'s required to heat 10 lb. of iron from 68°F. to 168°F.? (113)

42. 100 gm. of water at 10° C. are mixed with 200 gm. of water at 61° C. Assuming no heat lost, find the temperature of the mixture. (44°C.)

43. A brass rod is 1000 cm. long at 20°C. Find its expansion if the temperature is raised to 120°C. Coef. of linear expansion is 0.000020/C°. (2 cm.)

44. The volume of a given mass of gas is 91 cm.^{3} at S.T.P. Find its volume at 27°C. and 700 mm. of mercury. (109 cm.^{3})

45. Ethyl alcohol is kept in a glass container. The coefficient of linear expansion of glass is 0.000004/C°. The coefficient of absolute expansion of alcohol is 0.001110/C°. Find the coefficient of apparent expansion of alcohol in glass. (0.001098/C°)

46. The volume of a given mass of gas is 273 c.^{3} at 0°C. Find its volume at -10°C. (263 cm.^{3})

47. 250 calories of heat are supplied to 5 gm. of water at 17°C. Find the resulting temperature. (67°C.).

48. Find the thermal capacity of an aluminum calorimeter of mass 100 gm. (Specific heat of aluminum is 0.217 .) (21.7 cal.).

49. What minimum weight of water at 100°C. must be poured on a block of ice at 0°C. weighing 200 gm. wt. in order to melt it? (160 gm.)

50. How much ice at 0°C. must be mixed with 20 gm. of steam at 100°C. to produce a final temperature of 40°C.? (100 gm.)

51. The boiling point of oxygen is -183°C. What is its boiling point on the Fahrenheit scale? (-297.4)

52. What is the water equivalent of a copper calorimeter of mass 150 gm.? (14.25 gm.).

53. How much heat is required to melt 50 lb. of ice at 0°F.? (8000 B.t.u.)

54. An iron ball of mass 500 gm. at a temperature of 100°C. is dropped on a large block of ice at 0°C. What is the maximum mass of ice that could be melted? (75 gm.)

55. The latent heat of fusion of copper is 43 cal./gm. What is its heat of fusion in B.t.u./lb.? (7.74)

56. How much heat is required to melt 500 gm. of ice without changing its temperature? (40,000 cal.)

57. 440 gm. of metal at 450^{o}C. were placed into 600 gm. of water at 20^{o}C. The final temperature was 75^{o}C. Assuming no loss of heat, find the specific heat of the metal. (0.2)

58. The water equivalent of a calorimeter is 33 gm. and its specific heat is 0.22. Calculate the mass of the calorimeter. (150 gm.)

59. A calorimeter with water equivalent of 30 gm. contains 270 gm. of water at 40^{o}C. If 12 gm. of steam at 100^{o}C. are bubbled into it with no heat loss, find the final temperature. (63^{o}C.)

60. Convert 120 B.T.U. of heat to calories. (30,240)

61. In an experiment to determine the latent heat of vaporization of water the following data were obtained:

Wt. of empty calorimeter..................100 gm.
Sp. heat of calorimeter...................0.25
Wt. of calorimeter plus water at 15^{o}C.....300 gm.
Temperature of steam......................95^{o}C.
Wt. of calorimeter, water and condensed steam.................................307.5 gm.
Final temperature.........................35^{o}C.

From the recorded data calculate the latent heat of vaporization of water. (540 cal./gm.)

62. The following are observations on an experiment to determine the specific heat of aluminum:

Wt. of copper calorimeter.................200 gm.
Sp. heat of copper........................0.093
Wt. of water in calorimeter...............210 gm.
Temperature of water......................17^{o}C.
Wt. of aluminum...........................150 gm.
Temperature of aluminum...................98^{o}C.
Final temperature.........................27^{o}C.

Find the specific heat of aluminum. (0.215)

63. The following observations were made during an experiment:

Wt. of aluminum calorimeter...............60 gm.
Wt. of calorimeter and water..............170 gm.
Wt. of calorimeter, water and condensed steam.................................175 gm.
Initial temperature of water and calorimeter............................10^{o}C.
Final temperature.........................35^{o}C.
Water equivalent of calorimeter...........10 gm.

From these data calculate the heat of vaporization of water. (535 cal./gm.)

64. A bulb and stem of a mercury thermometer contain 0.6 cm.^{3} of mercury when the temperature is 0^{o}C. What is the distance between 0^{o}C. and 100^{o}C. marks if the internal diameter of the stem is 0.2 mm.? Coef. of apparent expansion of mercury in glass is 0.000156/C^{o}). (29.8 cm.)

65. To what temperature must two rods of copper and iron each 100 cm. long at 15^{O}C. be heated in order that they differ in length by 1 mm.? (Coef. of linear expansion of copper and iron are 0.000018 and 0.000012/C^{O} respectively.) (181.7^{O}C.)

66. The thermal capacity of 300 gm. of iron is 33.9 cal/C^{O}. What is the specific heat of iron? (0.113)

67. The density of air at S.T.P. is 1.29 gm/litre. What is its density at 68^{O}F. and 70 cm. of mercury? (1.1 gm./l.)

68. A glass flask of volume 1000 cm.^{3} will overflow when the temperature is raised to 50^{O}C.? The coefficient of linear expansion of glass is 0.0000092/C^{O}. The coef. of volume expansion of mercury is 0.000180/C^{O}. (4.57 cm.^{3})

69. 1000 gm. of boiling water at 100^{O}C. are poured into an aluminum tea kettle at 20^{O}C. The final temperature of the two is 86^{O} C. Assuming no heat losses, find:

(a) the water equivalent of the tea kettle. (212.1 gm)

(b) its mass (Sp. heat of Al. is 0.212 cal/gm/C^{O}.) (1000 gm.)

70. 200 gm. of copper at a temperature of 100^{O}C. are dropped into a calorimeter containing 140 gm. of water at 15^{O}C. The final temperature is 25^{O}C. Calculate the water equivalent of the calorimeter. (Sp. heat of Cu. is 0.10.) (10 gm.)

71. A copper calorimeter weighing 24 gm. contains 100 gm. of water at 19.5^{O}C. 12.8 gm. of ice at 0^{O}C. are added and the temperature drops of 8.5^{O}C. Calculate the latent heat of ice. (Sp. heat of copper is 0.10.) (79.5 cal/gm.)

72. The density of mercury at 20^{O}C. is 13.55 gm/cm.^{3} Find its density at 100^{O}C. Coefficient of volume expansion of mercury is 0.00018/C^{O}. (13.36 gm/cm.^{3})

73. Find the latent heat of ice from the following observations:

Mass of calorimeter........................60.0 gm.
Specific heat of calorimeter...............0.10
Mass of calorimeter and water..............170.0 gm.
Temperature of water.......................35.0^{O}C.
Mass of calorimeter, water and melted ice..211.0 gm.
Final temperature..........................5.0^{O}C.
Original temperature of ice................0^{O}C.

(79.9 cal/gm.)

SOUND

74. The time interval between the flash and the report of a gun is 44 seconds for a distance of 9.25 miles. Find the velocity of sound. (1110 ft/sec.)

75. The time interval between a flash of lightning and the clap of thunder was 11 seconds. Find the distance of the storm from the observer if the velocity of sound is 1120 ft./sec. (2.3 mi.)

76. The time interval between a flash of lightning and the calp of thunder was 5 sec. Find the distance of the storm from the observer is the velocity of sound is 1120 ft./sec. (5600 ft.)

77. The time interval between a flash of lightning and the clap of thunder is 5 seconds. The lightning appeared due south of the observer. A wind at 15 miles per hour was blowing from the north at this time. Find the distance of the storm from the observer if the velocity of sound is 1120 ft./sec. in still air. (5490 ft.)

78. An echo from the bottom of a sea was heard 1.2 seconds after the sound was produced at the surface. Find the depth of the sea if the velocity of sound in sea water is 4800 ft./sec. (2280 ft.)

79. Two notes of frequencies 256 and 260 are produced simultaneously. How many beats are heard? (4/sec.)

80. What is the wave length of a note produced by a body vibrating at 256 cycles per second if the velocity of sound in air is 34,000 cm./sec.? (133 cm.)

81. What is the frequency of the note produced by a disc rotating at 300 revolutions per minute if the disc has 50 uniformly spaced holes? (250/sec.)

82. A siren disc, having 30 evenly spaced teeth makes 200 revolutions per minute while a strip of metal is held against the teeth. If the velocity of sound in air is 1120 ft./sec., what is the wave length of the note produced? (11.2 ft.)

83. A siren disc having 60 holes is rotating uniformly 1000 times in 160 seconds.

(a) Find the frequency of the note produced. (375)

(b) Find the wave length if the velocity of sound in air is 345 m./sec. (0.92 m.)

84. A vibrating tuning fork is held over the mouth of a jar and the level of water adjusted until the jar is in resonance with the fork. The column of air is 10.0 in. long.

Calculate the frequency of the fork if the velocity of sound is 1080 ft./sec. (324/sec.)

85. A closed tube 26.9 cm. long is in resonance with a tuning fork of frequency 320 cycles per second. Find:

(a) the wave length of the note produced. (108 cm.)

(b) the velocity of sound in air. (344 m./sec.)

86. A closed pipe produces a note of frequency 380 cycles per second. When the pipe is shortened by 1.1 cm., the frequency of the note is 400 cycles per second. Assuming that in each case the fundamentals are produced, find

(a) the velocity of sound in air. (22 cm.)

(b) the original length of the pipe. (334 m/sec.)

87. A stretched string, of length 100 cm., vibrates at 80 cycles per second. What is the frequency of 50 cm. of the same wire under the same tension? (160/sec.)

88. A string 100 cm. long vibrates at 80 cycles per second. What is the frequency of the same string if its length is reduced to 50 cm. and the tension is multiplied by 4? (320/sec.)

89. A wire vibrates 100 cycles/sec. under a tension of 2 lb. wt. Under what tension will the same wire vibrate 150 cycles/sec.? (4.5 lb.wt.)

90. A wire, 1 meter long, vibrates at 288 cycles/second. A bridge is placed 40 cm. from one end of the wire. Find the frequencies of each of the two segments. (720 and 480 cycles/sec.)

91. A motorist passes a standing policeman at 60 miles per hour. The policeman blows a whistle having a frequency of 1400 cycles per second. What is the frequency of the note heard by the motorist as he travels away from the policeman? Velocity of sound is 1120 ft/sec. (1290/sec.)

92. A 220 yard dash was timed at 23.2 seconds. What would have been the time had the timer standing at the finish line listened to the sound instead of watching the flash of the starting gun? Temperature of air was 20°C. (22.6 sec.)

93. A string vibrates at 300 cycles per second under a tension of 4500 gm. wt. (450/sec.)

94. How many teeth must a saw have to produce a note having a frequency of 1350 cycles per second when rotated at 540 revolutions per minute? (150)

95. A string 30 cm. long vibrates at 540 cycles per second under a tension of 4000 gm. wt. Find its frequency under a tension of 16000 gm.wt. (1080)

96. A tuning fork having a frequency of 256 cycles per second is in resonance with a closed tube 13.0 inches long. Find the velocity of sound in air. (1109 ft./sec.)

97. A closed tube 40.0 cm. long is in resonance with a tuning fork having a frequency of 220 cycles per second. Find the temperature of the air. (33.3°C.)

98. In sounding a lake, the time between producing a sound and hearing the echo is 0.75 sec. If the velocity of sound in water is 4750 ft/sec., find the depth of the lake. (1781 ft.)

99. At what speed must an aircraft travel to break the sound barrier? (Approx. 750 mi/hr.)

100. A man sets his watch by the whistle of a factory 2,200 ft. away. By how many seconds is his watch out of time assuming that the speed of sound is 1100 ft./sec. (2.0 sec.)

Date Due
